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THE ONTARIO ASSESSMENT INSTRUMENT POOL

CHEMISTRY III

SENIOR DIVISION

Ministry of Education
Ontario

Hon. Bette Stephenson, M.D., Minister
Harry K. Fisher, Deputy Minister

CHEMISTRY III
SENIOR DIVISION
ONTARIO ASSESSMENT INSTRUMENT POOL

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Queen's Park
Toronto, Ontario

ISBN 0-7743-6790-3
ISSN 0228-0957
ON 02180

This is package number three (of four packages) of assessment materials related to the Ministry of Education curriculum guideline for Chemistry, Senior Division. It consists of a list of objectives for S-17 E, multiple-choice instruments, and statistical tables.

Cost: $15.00 per package

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The Ontario Assessment Instrument Pool (Chemistry) consists of four (4) packages of assessment materials related to the S-17D (1966) and S-17E (1967) Senior Division Chemistry Curricula. Additional materials will be published as they become available.

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Chapter 13

Multiple Choice Instruments

a) S-17E OBJECTIVES

UNIT 1

INTRODUCTION

At the conclusion of this unit, the student should be able to:

(1) define: scientific method, precision, accuracy, experimental error.

(2) distinguish between accuracy and precision.

(3) recognize that all measurements contain uncertainty.

(4) discuss the use of models in chemistry.

(5) graph experimental data.

(6) design and/or perform an experiment to test a hypothesis.

(7) calculate experimental uncertainty.

(8) recognize sources of experimental error.

UNIT 2

ENERGY EFFECTS IN CHEMICAL REACTIONS

At the conclusion of this unit, the student should be able to:

2.1.a(1) describe what is meant by the water-gas reaction.

a(2) write the thermochemical equations for the water-gas reaction (production & use).

a(3) write the thermochemical equation for the producer gas reaction.

a(4) define enthalpy of reaction (heat change of a reaction) in proper units.

a(5) define exothermic reaction and endothermic reaction.

a(6) identify exothermic and endothermic reactions from thermochemical equations.

a(7) tabulate enthalpy in the production and use of water gas.

a(8) recognize the industrial value of the producer gas-water gas cycle.

a(9) define thermochemistry.

2.1.b(1) define the following types of heats of reaction:

(i) heat of combustion
(ii) heat of dilution
(iii) heat of formation
(iv) heat of decomposition
(v) heat of ionization
(vi) heat of neutralization
(vii) heat of solution

b(2) define molar heat of formation (enthalpy) of a compound.

b(3) define various molar heats of reaction (solution, dilution, neutralization, etc.).

b(4) define bond energy of a molecule.

b(5) recognize that the source of energy lies in the making and breaking of bonds.

b(6) recognize that heat content of a substance is given in units of kilojoules per mole (kJ/mol).

2.1.c(1) state:

(i) Law of Additivity of Reaction Heats (Hess's Law of Constant Heat Summation)

(ii) Law of Conservation of Energy

c(2) tabulate experimental data in the Heat of Reaction experiment.

c(3) analyze data from the Heat of Reaction experiment.

c(4) deduce the enthalpy inter-relationship in the Heat of Reaction experiment.

c(5) represent energy changes in the Heat of Reaction experiment diagramatically.

c(6) generalize from the results of the Heat of Reaction experiment to similar situations.

c(7) apply the formula $Q = mc\triangle t$ to energy changes in the Heat of Reaction experiment.

c(8) define calorimetry.

2.1.d(1) design experimental procedures for the collection of data on heat changes.

d(2) describe an appropriate method of collecting experimental data in the Heat of Neutralization experiments.

d(3) recognize sources of error inherent in the collection of experimental data.

d(4) tabulate experimental data from the Heat of Neutralization experiment.

d(5) analyze data from the Heat of Neutralization experiment.

d(6) generalize from results of the Heat of Neutralization experiment to similar situations.

d(7) apply the formula $Q = mc\triangle t$ to energy changes in the Heat of Neutralization.

2.1.e(1) calculate the heats of reaction, $\triangle H$ in the Heat of Reaction and the Heat of Neutralization experiments.

e(2) calculate the heat of reaction, $\triangle H$, given a series of thermochemical equations.

e(3) derive $\triangle H$ values from thermochemical equations.

e(4) understand that the sign of the heat change is reversed when the direction of reaction is reversed.

e(5) predict heat of reaction from $\triangle H$ values.

e(6) calculate heats of reaction for a given amount of reactant from the molar heat of reaction.

e(7) analyze molar heats of combustion of fuels and decide which is better.

e(8) define "fuel value" of a substance.

e(9) be aware of the non-renewable nature of chemical sources of energy.

e(10) calculate heat of reaction given molar heats of formation.

e(11) calculate bond energy given a series of thermochemical equations.

e(12) use apparatus to obtain useful data in experiments.

e(13) perform experimental work showing awareness of sources of error.

e(14) design an experiment to measure the heat of combustion of a given solid.

e(15) design an experiment to determine which one of a group of fuels is the most efficient.

e(16) design experiments to measure and compare heat quantities involved in various reactions.

2.2.a(1) define kinetic and potential energy.

a(2) recognize chemical energy as one form of stored energy.

a(3) explain the relationship between gain in kinetic energy and loss in potential energy.

a(4) describe the importance of chemical sources of energy.

a(5) apply the relation $\triangle H = \triangle H$ products $- \triangle H$ reactants.

a(6) recognize that bond formation results in energy being stored in a molecule.

a(7) describe the kinetic energy of a gaseous molecule as translational, vibrational, and rotational.

a(8) **state** the 0 K can never be attained since electronic energy is always present.

2.2.b(1) **recognize** that energy is conserved in a chemical reaction.

b(2) **explain** why energy is conserved in a chemical reaction.

b(3) **construct** a Born-Haber cycle given a series of thermochemical equations.

b(4) **illustrate** the conservation of energy in a chemical reaction by means of an energy level diagram.

b(5) **explain** why a chemical reaction is endothermic in one direction if it is exothermic in the other direction.

b(6) **define:** $\triangle H$, $\triangle G$, and $\triangle S$.

b(7) **explain** how $\triangle S$ makes it possible to have spontaneous endothermic chemical reactions.

b(8) **use** $\triangle G = \triangle H - T\triangle S$ to predict whether a reaction will occur or not.

b(9) **recall** that the total energy of a system in a given state is always the same.

2.3.a(1) **define** bond energy.

a(2) **name** the various forms of energy stored in gaseous molecules.

a(3) **define** molar heat content of a molecule.

a(4) **be aware** of relative magnitudes of energies involved in phase changes, chemical changes and nuclear changes.

a(5) **relate** molar heat content to the energy of a molecule.

2.3.b(1) **describe** the various energy changes that occur as a solid is heated through to the gas phase.

b(2) **describe** the energy changes on warming a solid substance in terms of translational, vibrational, and rotational energies.

b(3) **explain** the absorption or evolution of energy during phase changes.

b(4) **relate** the absorption and evolution of energy during phase changes to changes in molecular rotational, vibrational and translational energy.

b(5) **recognize** the importance of bond energy as the major factor determining evolution or absorption of energy in a chemical reaction.

2.4.a(1) **recognize** the magnitude of the energies involved in nuclear reactions.

a(2) **define** fission reaction.

a(3) **define** fusion reaction.

a(4) **write** equations for fission and fusion reactions.

a(5) **recognize** that matter and energy are interconvertible ($E=mc^2$).

UNIT 3

THE RATES OF CHEMICAL REACTIONS

At the conclusion of this unit, the student should be able to:

3.1.a(1) **define** the term rate of reaction.

a(2) **use** appropriate units to express rate of reaction.

a(3) **recognize** that reaction rates must be determined experimentally.

3.2.a(1) **list** the factors which affect rate of reaction.

a(2) **recall** examples of slow and fast reactions.

3.2.b(1) **recall** that the rate of a reaction increases as the concentration of either reactant increases.

b(2) **explain**, using collision theory, the effect of concentration on rate.

b(3) **state** that surface area of solid reactants can often affect rate of reaction.

b(4) **explain**, using collision theory, the effect of surface area on rate.

b(5) **solve** problems involving the qualitative effect of concentration on reaction rate.

3.2.c(1) **define** the terms:

(i) reaction mechanism
(ii) rate-determining step
(iii) chain reaction

c(2) **identify** the rate-determining step in mechanisms of a non-chemical nature.

c(3) **solve** qualitative problems involving reaction mechanisms.

3.2.d(1) **state** information given by the rate equation for a reaction.

d.2 **define** the term reaction order.

d(3) **measure** reaction times for varying initial conditions.

d(4) **tabulate** data which illustrate the relation between rate and concentration.

d(5) **graph** data such as reactant concentration versus reaction time.

d(6) **deduce** the relationship between reaction rate and reactant concentration from experimental data.

d(7) **interpolate** and/or **extrapolate** observed concentration-time data.

3.2.e(1) **recall** that reaction rate increases as temperature increases.

e(2) **define** the term threshold energy.

e(3) **explain**, using collision theory, the effect of temperature on reaction rate.

e(4) **recognize** the concept of threshold energy in a non-chemical setting.

e(5) **draw** graphs to illustrate kinetic energy distribution.

e(6) **represent** the concept of threshold energy on a reaction potential energy diagram or on a kinetic energy distribution graph.

e(7) **use** kinetic energy distribution graphs on which threshold energy for a reaction is indicated:

(i) to show the effect of temperature on rate
(ii) to show the effect of a catalyst on rate
(iii) to compare the rates of slow and fast reactions

e(8) **tabulate** data which illustrate the relationship between rate and temperature.

e(9) **graph** data to illustrate clearly the relationship between rate and temperature.

e(10) **deduce** the relationship between reaction rate and temperature from experimental data.

e(11) **interpolate** between points to give a continuous graph of reaction time versus temperature and/or extrapolate the graph beyond actual observations.

3.3.a(1) **define** the terms:

(i) activation
(ii) activated complex

a(2) **draw** a potential energy diagram for a reaction or state in own words information from a given energy diagram.

a(3) use potential energy diagrams to illustrate

(i) the difference between slow and fast reactions
(ii) reactions that involve a series of steps

a(4) solve problems requiring an understanding of the concepts activation energy and activated complex.

3.3.b(1) recall the relationship between heat of reaction and activation energies for forward and reverse reactions.

b(2) translate given heats of reaction and activation energies into potential energy diagrams and vice-versa.

b(3) solve problems requiring use of the relationship between activation energies and heat of reaction.

3.3.c(1) define the terms:

(i) catalyst
(ii) catalysis

c(2) recall that catalysts provide different reaction paths with lower activation energies.

c(3) recognize the concept of activation energy lowering in a non-chemical setting.

c(4) represent catalysis on the potential energy diagram for a reaction.

c(5) describe the observed effect of a catalyst.

3.3.d(1) recall that catalysis alters the mechanism of a reaction and that the catalyst is regenerated in one of the later steps.

d(2) recall examples to illustrate the stepwise nature of catalysis.

d.(3) recognize catalysis in living organisms.

UNIT 4

EQUILIBRIUM

At the conclusion of this unit, the student should be able to:

4.1.a(1) **define** the equilibrium state in terms of constant macroscopic properties.

a(2) **recall** chemical equilibrium as a dynamic process.

a(3) **distinguish** between equilibrium and steady state systems.

a(4) **identify** a described system as steady state or equilibrium.

4.2.a(1) **distinguish** between macroscopic and microscopic properties.

b(2) **identify** macroscopic and microscopic properties of an equilibrium system.

b(3) **state**, with reasons, whether or not a described system is at equilibrium.

4.2.b(1) **define** each of the following in terms of the molecular processes involved:

(i) saturation with respect to solubility
(ii) saturation with respect to liquid/vapour systems
(iii) chemical equilibrium systems

4.2.c(1) **state** the relationship between the rates of the forward and reverse reactions of an equilibrium system.

c(2) **recognize** that the equilibrium composition of a chemical system is not dependent upon the direction from which equilibrium is approached.

4.2.d(1) **explain** why altering each of the following conditions alters the equilibrium composition:

(i) concentration of reactants and products
(ii) temperature

d(2) **explain** the effect of catalysis on equilibrium.

4.2.e(1) **differentiate** between rate of reaction and extent of reaction.

4.2.f(1) **state** LeChatelier's Principle.

f(2) **apply** LeChatelier's Principle to predict the effect on the equilibrium composition in a given physical or chemical system of changes in:

(i) concentration
(ii) pressure
(iii) temperature
(iv) volume
(v) non-reacting species

4.2.g(1) **apply** LeChatelier's Principle to select the experimental conditions that would theoretically give the best yield of products for a given chemical reaction.

4.3.a(1) **define** equilibrium constant.

a(2) **predict** whether reactants or products are favoured at equilibrium, given the equation and the equilibrium constant value for a given reaction.

4.3.b(1) **write** the equilibrium constant expression, given the equation for a chemical reaction.

b(2) **write** the equilibrium constant expression, given equations containing pure solids and liquids.

b(3) **apply** LeChatelier's Principle to determine the effect on the equilibrium constant value of given systems for changes in:

(i) temperature
(ii) pressure
(iii) concentration

4.3.c(1) **show** that, for any chemical reaction, the equilibrium constant for the reverse reaction is the inverse of the equilibrium constant for the forward reaction.

4.3.d(1) **define** entropy

d(2) **state** the two factors which determine the equilibrium composition.

d(3) **state** which factor determining the equilibrium composition is more important at high and low temperatures.

d(4) **select** between varying degrees of randomness.

d(5) **predict** whether the forward or reverse reaction in a given example is favoured by the drive towards:

(i) minimum energy
(ii) maximum randomness (entropy)

4.3.e(1) **calculate** the equilibrium constant for a chemical reaction given the equation and equilibrium concentrations of reactants and products.

e(2) **calculate** the equilibrium constant for a chemical reaction given the equation and initial concentrations of reactants and one equilibrium concentration of either reactants or products.

e(3) **calculate** the value of the equilibrium concentration of one substance given the value of the equilibrium constant and the remaining equilibrium concentrations.

e(4) **calculate** the value of the equilibrium concentrations of reactants and products given the initial concentrations of reactants and the equilibrium constant.

4.3.f(1) **perform** an experiment to illustrate the law of chemical equilibrium.

UNIT 5

SOLUBILITY EQUILIBRIA

At the conclusion of this unit, the student should be able to:

5.1.a(1) define solubility.

a(2) define saturated solution.

a(3) recognize that solubility is a special case of equilibrium.

a(4) recognize, on a macroscopic level, when a saturated solution has been formed.

5.1.b(1) list the factors affecting the rate of dissolving of a solid.

b(2) explain why various factors affect the rate of dissolving of a solid.

b(3) define precipitation.

b(4) explain why concentration of solution affects rate of precipitation of a solid but concentration of solution does not affect the rate of dissolving of a solid.

b(5) discuss the factors affecting the rate of precipitation.

b(6) recognize that solubility equilibrium is dynamic.

b(7) explain on a molecular level, the dynamic nature of equilibrium.

5.1.c(1) define entropy (randomness).

c(2) explain, using entropy and energy, why some solids are more soluble in one solvent than in another.

c(3) explain why temperature changes always affect the solubility of a solid.

5.1.d(1) discuss the factors affecting the solubility of gases in liquids.

d(2) explain why entropy and energy changes always affect the solubility of gases in liquids.

5.2.a(1) define:

(i) electrolyte
(ii) nonelectrolyte

a(2) give examples of electrolytes and non-electrolytes.

5.2.b(1) identify types of compounds which are electrolytes.

b(2) differentiate between ionic covalent compounds.

b(3) write and balance net ionic equations for precipitation reactions.

5.2.c(1) state the general rules for solubility of compounds in water.

c(2) predict qualitatively whether a given salt is soluble in water.

5.2.d(1) define solubility-product constant K_{sp}.

d(2) write the solubility-product constant (K_{sp}) for a given salt.

d(3) relate the magnitude of K_{sp} to the solubility of a given substance.

d(4) design and/or perform an experiment to determine the K_{sp} for a given compound.

5.2.e(1) calculate the solubility of salts in required concentration units, given the K_{sp}.

5.2.f(1) predict qualitatively if a precipitate will form given a mixture of ions in aqueous solution.

f(2) predict quantitatively, if a precipitate will form, given concentrations of the reacting substances and the K_{sp}.

5.2.g(1) design methods to separate specific salts from an aqueous solution of ions.

5.2.h(1) solve simple problems to determine K_{sp} for a given substance.

UNIT 6

AQUEOUS ACIDS AND BASES

At the conclusion of this unit, the student should be able to:

6.1.a(1) **define** the terms strong and weak electrolytes.

a(2) **describe** an experiment to identify solutions as being either strong or weak electrolytes.

a(3) **distinguish** between strength and concentration of electrolytes.

a(4) **design** and/or **perform** an experiment to distinguish between strong and weak electrolytes.

6.1.b(1) **recognize** that water is a very weak electrolyte.

b(2) **identify** the ions formed by the ionization of water.

b(3) **write** the equation for the ionization of water.

b(4) **recall** the numerical value of the ion product constant for water, K_w, at $25^{\circ}C$.

b(5) **state** the effect of temperature on K_w.

6.1.c(1) **recognize** that $[H^+(aq)] = [OH^-(aq)] = 10^{-7}$ mol/L for pure water.

c(2) **apply** LeChatelier's Principle to an explanation of the shift in the equilibrium concentrations of $H^+(aq)$ and $OH^-(aq)$ when either $H^+(aq)$ or $OH^-(aq)$ are added to water.

c(3) **calculate** the $[H^+(aq)]$ and $[OH^-(aq)]$ in an aqueous solution given the mass of a strong acid or base and the total volume of the solution.

6.2.ab(1) **distinguish** between an operational and a conceptual definition.

ab(2) **state** an operational definition of:

(i) an acid
(ii) a base

ab(3) **state** a conceptual definition of:

(i) an acid
(ii) a base

ab(4) distinguish between operational and conceptual definitions of an acid and a base.

ab(5) identify the acid and base in a reaction given the chemical equation for the acid-base reaction.

ab(6) state the colour of various indicators in acidic and basic solutions.

6.2.c(1) write the formula of:

(i) a hydrated proton
(ii) the hydronium ion

6.2.d(1) define:

(i) acid-base titration
(ii) neutralization
(iii) acid-base indicator
(iv) end-point
(v) equivalence point

d(2) calculate at any point in an acid-base titration the $[H^+(aq)]$ and $[OH^-(aq)]$.

d(3) solve quantitative acid-base titration problems.

d(4) write the formulae for the products of a given acid-base reaction.

d(5) demonstrate the usage of pipettes and burettes.

d(6) state to the proper accuracy the volume measured in an acid-base titration using a burette and a pipette.

d(7) design an experiment to determine the concentration of a given acidic or basic solution.

d(8) perform an acid-base titration.

d(9) tabulate data obtained in a titration experiment.

d(10) select the appropriate indicator for specific acid-base titration.

d(11) design and/or perform an experiment to determine the relative acidity of common household substances.

6.2.e(1) define pH.

e(2) define pOH, pK.

e(3) state whether a solution is acidic, basic or neutral, given the pH.

e(4) **calculate** the $[H^+(aq)]$ or $[OH^-(aq)]$ of a solution given:

(i) an integral pH value
(ii) any pH value

e(5) **calculate** the pH of a solution given:

(i) the $[H^+(aq)]$ or $[OH^-(aq)]$

(ii) the $[H^+(aq)]$ or $[OH^-(aq)]$ as a power-of-ten.

6.3.a(1) **write** the K_a expression for the ionization of a monoprotic acid.

a(2) **define:**

(i) monoprotic acid
(ii) polyprotic acid

a(3) **write** the chemical equation(s) for each step in the ionization of a polyprotic acid.

a(4) **define** K_a

a(5) **write** the K_a expression for each ionization reaction.

a(6) **order** acids according to strengths using their K_a value.

a(7) **calculate** the K_a value of an acid from experimental data.

a(8) **calculate** the $[H^+(aq)]$ in an acid solution given the K_a value and the initial acid concentration.

a(9) **calculate** the $[H^+(aq)]$ of a solution containing both an acid and its salt given the K_a value and experimental data.

a(10) **recognize** indicators as weak acids or bases.

a(11) **perform** an experiment to determine the K_a value of a weak acid.

a(12) **apply** LeChatelier's principle to acid-base reactions.

6.3.b(1) **state** the proton transfer (Bronsted-Lowry) definitions of acids and bases.

b(2) **recognize** that the reaction between an acid and a base produces another acid and another base.

b(3) **identify** each reactant and product as either a Bronsted-Lowry acid or base.

b(4) **recognize** the ionization equation of an acid as a proton transfer (Bronsted-Lowry) reaction.

b(5) **predict** whether products or reactants are favoured in a Bronsted-Lowry system given K_a values.

b(6) **write** the equations for the reactions between Bronsted-Lowry acids and bases.

6.3.c(1) **state** the Lewis definition of an acid and base.

c(2) **identify** the Lewis acid and base given the equation for the reaction.

UNIT 7

OXIDATION - REDUCTION REACTIONS

At the conclusion of this unit, the student should be able to:

7.1.a(1) define the following terms:

(i) anode
(ii) cathode
(iii) external circuit
(iv) electrolyte
(v) oxidation
(vi) reduction
(vii) oxidizing agent
(viii) reducing agent
(ix) half-cell reaction
(x) electrochemical cell
(xi) salt bridge

a(2) write half-cell oxidation reactions.

a(3) write half-cell reduction reaction.

a(4) state the purpose of the salt bridge.

a(5) identify a reaction as redox or non-redox.

7.1.b(1) describe the relevance of the electrochemical series.

b(2) write redox equations using the electrochemical series.

b(3) write net ionic redox equations using the electrochemical series.

b(4) recognize that redox reactions can occur without electrodes.

b(5) express a redox reaction in terms of oxidation and reduction half reactions.

7.1.c(1) identify the following parts of electrochemical and/or electrolytic cells:

(i) anode
(ii) cathode
(iii) external circuit
(iv) electrolyte
(v) oxidation
(vi) reduction
(vii) oxidizing agent
(viii) reducing agent
(ix) half-cell
(x) salt bridge
(xi) direction of electron flow
(xii) direction of ion flow

7.2.a(1) perform an experiment to determine the cell potential of an electrochemical cell.

a(2) write half-cell reactions from experimental data.

a(3) recognize the hydrogen half-cell as an arbitrary standard for E^{o} values

a(4) recall that a half-cell potential cannot be measured in isolation.

a(5) recognize that an item to be electroplated should be the cathode of an electrolytic cell.

7.2.b(1) recognize that E^{o} values are a measure of the tendency for a reaction to occur.

b(2) calculate the overall cell potential from E^{o} values.

b(3) predict spontaneity of reactions given a table of half-cell potentials.

b(4) predict the qualitative effect of changes in electrolyte concentrations on E^{o} values.

b(5) use E^{o} values to order substances according to their tendencies to be reduced or oxidized.

b(6) design an experiment to determine an E^{o} potential for a particular half-cell reaction.

7.2.c(1) state a set of rules for the assignment of oxidation numbers.

c(2) determine the oxidation number of any element in a compound, ion, or substance.

c(3) identify oxidation as an increase in oxidation number.

c(4) **identify** reduction as a decrease in oxidation number.

7.3.a(1) **balance** a redox reaction using half-reactions.

a(2) **balance** a redox reaction using oxidation numbers.

7.4.a(1) **define:**

(i) electrolysis
(ii) electrolytic cell

a(2) **recognize** electrolysis as a redox reaction.

a(3) **write** equations for the electrolysis of:

(i) water
(ii) aqueous electrolytes
(iii) molten ionic compounds

7.4.b(1) **recognize** that electrolysis is a reverse reaction to that occurring in an electrochemical cell.

b(2) **design** an experiment to identify the products of electrolysis reactions.

b(3) **predict** overall cell potential of cells coupled in series.

b(4) **design** an experiment to measure the overall cell potential of cells coupled in series.

UNIT 8

EXPERIMENTAL BASIS FOR THE ATOMIC THEORY

At the end of this unit, the student should be able to:

8.1.a(1) state the law of conservation of mass.

a(2) state the law of definite proportions.

a(3) state the law of multiple proportions.

a(4) organize data from experiments to illustrate the law of definite proportions.

a(5) interpret experimental data illustrating the law of definite proportions.

a(6) interpret experimental data illustrating the law of multiple proportions.

a(7) compare predicted outcomes for percentage composition of a substance, based on a chemical formula, with that achieved by experiment.

a(8) design and/or perform experiments to determine empirical formulas.

a(9) determine empirical formula from experimental data.

a(10) recognize the need for a theory to explain the law of definite proportions.

a(11) recognize the need for a theory to explain the law of multiple proportions.

a(12) relate atomic theory to the law of definite proportions.

a(13) relate atomic theory to the law of multiple proportions.

a(14) solve problems illustrating the law of definite proportions.

a(15) solve problems illustrating the law of multiple proportions.

a(16) discuss the historical development of the atomic theory.

a(17) compare various models for the structure of matter.

8.1.b(1) **state** the law of combining gas volumes.

b(2) **state** Avogadro's hypothesis.

b(3) **propose** explanations for the law of combining gas volumes.

b(4) **solve** stoichiometric problems involving gas volumes.

b(5) **design and/or perform** experiments to illustrate the law of combining gas volumes.

b(6) **interpret** data gained from gaseous volume experiments.

8.1.c(1) **state** Faraday's laws of electrolysis.

c(2) **define:**

(i) a Faraday
(ii) a Coulomb

c(3) **relate** atomic theory to the results of Faraday's electrolysis experiments.

c(4) **design and/or perform** an experiment to illustrate Faraday's laws.

c(5) **solve** problems based on Faraday's laws.

8.2.a(1) **define** the term cathode ray.

a(2) **recognize** the conditions under which cathode rays are produced.

a(3) **state** the effects of magnetic field, electric field, and potential difference on cathode rays.

a(4) **describe** qualitatively the Thomson $\frac{e}{m}$ experiment.

a(5) **describe** qualitatively the Millikan oil drop experiment.

a(6) **solve** for "m" of an electron given "e" or vice versa.

a(7) **recognize** that there is an $\frac{e}{m}$ for positive ions.

8.2.b(1) **describe** qualitatively the operation of a mass spectrograph.

b(2) **solve** for "m" of a positive ion given "e" or vice versa.

b(3) **describe** qualitatively the Rutherford gold foil experiment.

b(4) **compare** the Rutherford and Thomson models of the atom.

8.3.a(1) **recognize** that light can be characterized by its frequency and/or wavelength.

a(2) **identify** the relative spectral positions of radio microwave, IR, visible, UV, and x-rays.

a(3) **solve** problems relating frequency and wavelength.

8.3.b(1) **recall** that the scattering of x-rays by a crystal can be used to reveal the spacing of atoms in that crystal.

b(2) **recognize** that molecules have natural frequencies of rotation and vibration.

b(3) **recognize** that molecules absorb radiation of specific frequencies dependent upon the type of bonds present.

b(4) **recall** that I.R. and M.V. spectrometry is used to give information on bond vibration and rotations.

UNIT 9

ELECTRON ARRANGEMENT AND THE PERIODIC TABLE

At the conclusion of this unit, the student should be able to:

9.1.a(1) state the equation relating the frequency, wavelength and velocity of light (universal wave equation).

a(2) state the speed of light in a vacuum in metres per second.

a(3) calculate frequency or wavelength from the equation

$$\lambda = \frac{c}{\nu}$$

a(4) define photon (quantum).

a(5) list the following radiation in order of increasing energy - infrared, red, ultraviolet, violet.

a(6) describe the wave-particle duality of light.

a(7) state Planck's equation ($E = h\nu$).

a(8) design and/or perform an experiment to examine emission spectra.

a(9) interpret data from emission spectra experiments.

9.1.b(1) distinguish between:

(i) continuous and line spectra
(ii) absorption and line spectra

9.1.c(1) solve problems using the Rydberg equation.

c(2) solve problems using Planck's equation.

9.1.d(1) relate the line spectrum of hydrogen to the existence of energy levels in the model of the atom.

d(2) recognize the existence of stationary states in the model of a hydrogen atom as proposed by Bohr.

d(3) calculate the energy of transition given the energy value for each level.

d(4) predict in which region of the electromagnetic spectrum emitted light of a given energy will appear.

9.1.e(1) define quantum number in terms of stationary states (energy levels) of an atom.

e(2) relate the value of n, principal quantum number, to the value of the energy level in the hydrogen atom model.

e(3) state the symbol, name, and mathematical restrictions for each of the 4 quantum numbers.

e(4) state the information about the location of an electron in an atom that each quantum number provides.

9.1.f(1) define orbital.

f(2) recognize that the term orbital refers to a description of the probable location of an electron.

f(3) write the symbols of the first 4 quantum numbers.

f(4) discuss the historical development of the wave-mechanical model of the atom.

9.1.g(1) describe the s and p orbital shapes.

9.1.h(1) calculate the number of

(i) orbital types
(ii) orbitals
(iii) electrons

allowed for any value of n.

9.2.a(1) list the orbital types in order of increasing energy for the same value of n.

a(2) recognize the difference in the energy level diagram for a hydrogen atom and that for a many electron atom.

9.2.b(1) state the Pauli Exclusion Principle.

b(2) state Hund's Rule.

b(3) apply the Aufbau Principle to list the order of fitting of orbital types.

b(4) draw representations to illustrate the electron arrangement for elements.

b(5) write electronic configurations for elements.

b(6) recognize that the similarity in the properties of elements in the same family can be explained by their electron configurations.

b(7) relate the arrangement of the periodic table to energy levels.

b(8) **identify** the position of the s,p, d and f blocks in the periodic table.

b(9) **state** that periods (horizontal rows) in the periodic table result from filling orbitals of approximately the same energy.

9.3.a(1) **define:**

(i) ionization energy
(ii) valence electron

9.3.b(1) **explain** why the second ionization energy of an element is always greater than its first ionization energy.

b(2) **describe** the trends in ionization energy in the periodic table:

(i) across a row
(ii) down a column

b(3) **explain** the trends in ionization energy in the periodic table

(i) across a row
(ii) down a column

b(4) **draw** representations to illustrate the electron arrangement of ions.

b(5) **state** the number of valence electrons possessed by an atom given:

(i) the electron configuration
(ii) a series of successive ionization energies
(iii) the atomic number

b(6) **state** that the transition elements are formed by filling d orbits of the penultimate energy level.

UNIT 10

MOLECULES IN THE GAS PHASE

At the conclusion of this unit, the student should be able to:

10.1.a(1) **define** the terms:

(i) chemical bonding
(ii) covalent bonding
(iii) covalent bond

a(2) **recognize** that all chemical bonding involves electrical attraction.

a(3) **recognize** that all chemical bonds form because electrons are simultaneously near two positive nuclei.

a(4) **recall** that energy is released when a chemical bond is formed.

a(5) **explain** why, in forming a chemical bond, that the energy is lower when the atoms are near each other.

a(6) **recognize** that the noble gases are monatomic.

a(7) **explain** why there is no covalent bond formation between atoms of the noble gases.

a(8) **draw**, for simple molecules:

(i) Lewis electron-dot diagrams
(ii) structural formulae
(iii) orbital representation

a(9) **define** electron affinity.

a(10) **relate** electron affinity to the formation of a chemical bond between atoms of nonmetals.

10.2.a(1) **explain** bonding capacity in terms of orbital occupancy.

a(2) **relate** valence to bonding capacity.

a(3) **predict** the bonding capacity of simple atoms by referring to their electronic arrangements.

a(4) **predict** the valence of simple atoms by referring to their electronic arrangements.

10.3.a(1) **define** the terms:

(i) ionic bond
(ii) ionic bonding

a(2) **explain** ionic bonding in terms of ionization energy differences.

a(3) **contrast** the ionic bond and covalent bond.

a(4) **compare** compounds (e.g. fluorine) of the second row elements as to:

(i) decreasing ionic character
(ii) increasing covalent character

a(5) **explain** why there is a graduation from an ionic bond in gaseous lithium fluoride to a pure covalent bond in gaseous fluorine.

a(6) **define:**

(i) electric dipole
(ii) polar molecule
(iii) polarity

a(7) **draw** a representation for the electric dipole of a given simple gaseous molecule.

a(8) **discuss** the effect of electric dipoles on bond strength.

a(9) **explain** the difference between polar molecules and nonpolar molecules.

a(10) **identify** the electric dipole in polar molecules.

a(11) **explain** why there is a graduation in ionic character between hydrogen and other atoms bonded to it.

10.4.a(1) **explain** the term molecular architecture (stereochemistry).

a(2) **relate** molecular shapes of simple molecules to orbital representations.

a(3) **identify** molecular shapes and bonding orbitals of simple molecules.

a(4) **recognize** that a molecular dipole will be determined by the sum of the dipoles of all the bonds.

a(5) **describe** the effect of electric dipoles on molecular shape.

a(6) **construct** models to represent various molecular shapes.

10.5.a(1) **define:**

(i) single bond
(ii) double bond
(iii) triple bond

a(2) **represent** the double bond in simple molecules using:

(i) Lewis electron-dot diagrams.
(ii) structural formulae.
(iii) orbital representations.

a(3) **explain** why a multiple covalent bond is in general stronger than a single covalent bond between the same atoms.

a(4) **define** cis-trans isomerism.

a(5) **show** why double bonds sometimes lead to cis-trans isomerism.

a(6) **design** and/or **perform** an experiment to investigate some of the properties of a pair of cis-trans isomers: e.g., maleic and fumaric acids.

UNIT 11

THE BONDING IN SOLIDS AND LIQUIDS

At the conclusion of this unit, the student should be able to:

11.1.a(1) **define** Van der Waals' forces.

a(2) **locate** (on the periodic table) those elements forming Van der Waals' liquids and solids.

11.1.b(1) **explain** why carbon forms a covalent network solid.

b(2) **define** the term allotrope.

b(3) **compare** the structure of diamond and graphite.

b(4) **state** some properties of the allotropic forms of carbon and relate these properties to the bonding and the packing of the atoms.

b(5) **locate** on the periodic table, those elements that form covalent network solids.

11.1.c(1) **explain** metallic bonding in terms of orbital occupancy and ionization energy.

c(2) **recall** the characteristic properties of metals.

c(3) **locate** on the periodic table, the metallic elements.

c(4) **explain** the properties of metals.

c(5) **relate** the trends in metallic characteristics on the periodic table to orbital occupancy and ionization energy.

c(6) **relate** the properties of Boron to the properties of metals and non-metals.

c(7) **explain** why some elements display metallic and non-metallic properties.

11.2(1) **construct** models to show the packing of atoms or ions in crystals.

11.2(2) **identify**, from models, the packing arrangement of atoms or ions in crystals.

11.2(3) **design** and/or **perform** an experiment to classify an unknown substance according to bonding type.

11.2.a(1) **explain** how the magnitude of Van der Waals' forces is affected by:

(i) number of electrons in the outermost atoms of a molecule
(ii) molecular size
(iii) molecular shape

11.2.b(1) **describe** the different forms of covalent network solid compounds.

11.2.c(1) **define** the term alloy.

c(2) **explain** why the properties of alloys are different from those of their components.

11.2.d(1) **define** lattice energy of a crystal.

d(2) **explain** the formation of ionic crystals.

d(3) **recall** the properties of ionic solids.

d(4) **compare** the conductivity of molten ionic substances and metals.

d(5) **compare** the mechanism by which electricity is conducted in both ionic substances and metals.

11.2.e(1) **define** the term polar molecule.

e(2) **explain** the solvent action of polar and non-polar substances.

e(3) **define** hydration.

e(4) **explain** the solvent action of water.

11.2.f(1) **state** the requirements for hydrogen bond formation.

f(2) **compare** the strengths of hydrogen bonding.

f(3) **explain** the effect of hydrogen bonding on the boiling points of liquids.

UNIT 12

CHEMISTRY OF CARBON COMPOUNDS

At the conclusion of this unit, the student should be able to:

12.1.a(1) **define** organic chemistry, organic compound, cracking, fractional distillation, petrochemical.

a(2) **identify** sources of organic compounds.

a(3) **recognize** various synthetic materials around home and school as being synthesized from organic substances.

a(4) **recognize** importance of coal for a fuel.

a(5) **recognize** petrochemical industries to be historically recent.

a(6) **dicuss** socioeconomic implications of rapid consumption of organic reserves.

a(7) **discuss** the ecological implications of the petrochemical industry (pollution, health).

a(8) **discuss** the importance of the petrochemical industry to modern society.

a(9) **recognize** the multiplicity of carbon compounds.

12.2.a(1) **define** empirical formula, molecular formula, structural isomerism, intermolecular, intramolecular forces.

a(2) **describe** bonding of elements in organic compounds.

a(3) **explain** by referring to bonding, the multiplicity of carbon compounds.

a(4) **identify** examples of intermolecular and intramolecular forces in carbon compounds.

a(5) **relate** properties of carbon compounds to bond strength.

a(6) **draw** for organic compounds:

(i) structural diagrams given their molecular formulae.
(ii) electron dot diagrams given their molecular formulae.
(iii) structural isomers given their molecular formulae.

a(7) **relate** melting points and boiling points to molecular size.

a(8) construct and/or identify models of carbon compounds.

12.2.b(1) identify products of combustion of hydrocarbons.

b(2) determine the empirical (simplest) formulae of carbon compounds, given experimental data.

b(3) identify molecular structures from experimental data.

b(4) design and/or perform an experiment to determine:

(i) empirical formulae
(ii) molecular formulae of carbon compounds
(iii) molecular structure of carbon compounds

12.2.c(1) identify alkyl group.

c(2) recognize the alkyl group as a structural entity.

12.3.a(1) define alkyl halide.

a(2) write:

(i) molecular formulae for alkyl halides
(ii) structural formulae for alkyl halides

a(3) compare the properties of the alkyl halides with reference to the alkyl group.

12.3.b(1) apply the terms

(i) oxidation
(ii) reduction
(iii) oxidizing agent
(iv) reducing agent
(v) oxidation number
(vi) half reaction

to reactions of carbon compounds.

b(2) recognize oxidation (reduction) in organic reactions.

b(3) write balanced equations for organic reactions.

b(4) discuss the economic and industrial importance of oxidation of organic compounds.

12.3.c(1) define functional group.

c(2) recognize R as representing an alkyl group in structural diagrams.

c(3) recall general structural formulae for functional groups.

c(4) **identify** functional groups given structural diagrams or models.

c(5) **write** structural formulae for functional groups given names or models.

c(6) **write** equations for synthesis of carbon compounds containing functional groups.

c(7) **relate** properties to functional groups.

c(8) **design** and/or **perform** an experiment to identify functional groups.

c(9) **design** and/or **perform** an experiment to synthesize organic compounds containing various functional groups.

12.4.a(1) **name** organic compounds using:

(i) IUPAC rules.
(ii) trivial.

a(2) **write** formulae for organic compounds.

12.5.a(1) **define** hydrocarbon, saturated hydrocarbon, unsaturated hydrocarbon, alkane, homologous series.

a(2) **recognize** alkanes from their formulae or models

a(3) **identify** straight-chained (unbranched), cyclic and branched alkanes, given their structural formulae or models.

a(4) **write** equations for reactions involving alkanes.

a(5) **describe** methods of refining alkanes.

12.5.b(1) **define** alkene, alkyne, addition reaction, substitution reaction.

b(2) **recognize** alkenes and alkynes from their structural formulae or models.

b(3) **identify** alkenes and alkynes, given their structural formulae or models.

b(4) **write** equations for reactions involving alkenes and alkynes.

b(5) **design** and/or **perform** experiments to identify alkanes, alkenes, and alkynes.

12.5.c(1) **define** aromatic compound.

c(2) **recognize** various symbolic representations of benzene and/or its derivatives.

c(3) **write** equations for reactions involving benzene and/or its derivatives.

c(4) **discuss** the stability of the benzene ring.

c(5) **discuss** the health hazards related to the use of benzene and its derivatives.

12.6.a(1) **define** polymer, monomer, polymerization, addition polymerization, condensation polymerization.

a(2) **recognize** that polymers have very large molecular weights.

a(3) **discuss** the growth of the polymer industry.

a(4) **write** structural formulae for simple polymers.

a(5) **write** equations for polymerization reactions.

a(6) **recognize** the widespread use of polymers in home and industry.

12.6.b(1) **define** polyamide, polyester.

b(2) **recognize** polyamides and polyesters from their structural formulae.

b(3) **write** structural formulae for polyamides (nylon) and polyesters.

b(4) **write** equations for the synthesis of polyamides and polyesters.

12.6.c(1) **define** protein, amino acid, peptide.

c(2) **recognize** protein as a polyamide.

c(3) **discuss** the economic and medical importance of proteins.

UNIT 13

THE THIRD ROW OF THE PERIODIC TABLE

At the conclusion of this unit, the student should be able to:

13.1.a(1) **identify** the elements of the third row.

a(2) **recall** the appearance of the third-row elements.

a(3) **classify** third-row elements as metals or non-metals.

a(4) **explain** the variation in physical properties of the third-row elements.

a(5) **solve** problems concerning the trends in properties of the third-row metals.

13.1.b(1) **explain** the physical properties of silicon.

b(2) **solve** problems relating the properties of silicon to those of other third-row elements

13.1.c(1) **explain** the physical properties of phosphorus, sulfur, chlorine and argon

c(2) **solve** problems relating the physical properties of phosphorus, sulfur, chlorine and argon to those of other third-row elements.

13.2(1) **compare** the third-row elements with respect to oxidizing (reducing) ability.

13.2.a(1) **explain** the relatively low reduction potentials for sodium, magnesium and aluminum.

a(2) **solve** problems concerning reactions in which third-row metals are reduced.

13.2.b(1) **explain** the intermediate oxidizing and reducing strength of silicon, phosphorus and sulfur.

b(2) **solve** problems concerning oxidation-reduction reactions involving silicon, phosphorus and sulfur.

13.2.c(1) **explain** the strength of chlorine as an oxidizing agent.

c(2) **solve** problems concerning oxidation-reduction reactions involving chlorine.

13.3(1) **recall** the acidic and basic properties of the hydroxides of the third-row elements.

13.3(2) **design** and/or **perform** an experiment to compare the hydroxides of third-row elements as acidic, basic or amphoteric.

13.3(3) **write** equations for acid/base reactions of the third-row elements.

13.3(4) **explain** the behaviour of the third-row hydroxides as acidic, basic or amphoteric.

13.3.a(1) **recall** the physical properties of NaOH and $Mg(OH)_2$.

13.3.b(1) **define** amphoterism.

13.3.c(1) **explain** the acidic properties of hydrated silicon dioxide.

c(2) **recall**

(i) the names
(ii) the structures
(iii) the relative strengths

of the oxyacids of phosphorus, sulfur and chlorine.

13.3. **compare** the observed acid-base behaviours of the third-row hydroxides to that expected from consideration of the ionization energies of the elements.

13.4.a(1) **recall** the occurrence in nature of the third-row elements.

a(2) **explain** the occurrence in nature of the third-row elements.

13.4.b(1) **recall** the methods of preparation of the third-row elements.

13.5(1) **translate** trends in properties of the third-row elements from tabular into graphic or verbal form.

UNIT 14

THE FOURTH-ROW TRANSITION ELEMENTS.

At the conclusion of this unit, the student should be able to:

14.1.a(1) **locate** the transition elements in the periodic table.

a(2) **recognize** the transition elements by their electron configuration.

a(3) **write** the electron configurations of the fourth row transition elements.

14.1.b(1) **recall** that all the transition elements are metallic.

b(2) **list** properties common to the transition elements.

b(3) **recognize** the range of reactivities of the transition elements.

b(4) **recognize** that transition elements form compounds and solutions which often have characteristic colours.

b(5) **recognize** that most transition elements have several possible oxidation states.

14.2.a(1) **recall** that transition elements form a great variety of complex ions.

a(2) **define** the term:

(i) complex ion
(ii) coordination ion
(iii) nearest neighbour

a(3) **define:**

(i) complex salt
(ii) double salt

a(4) **perform** an experiment to prepare:

(i) a complex salt
(ii) a double salt

14.2.b(1) **represent** the structures of complexes by diagrams and/or models.

14.2.c(1) **recognize** that the bonding in complex ions is mainly covalent.

c(2) **recognize** that d orbitals are not the only orbitals involved in bonding of complex ions.

c(3) **identify** the shapes of complexes given the orbitals involved in bonding, namely: sp - linear
dsp^2 - square planar
d^2sp^3 - octahedral
sp^3 - tetrahedral

14.2.d(1) **write** equations to illustrate the amphoteric behaviour of transition element complexes.

14.3.a(1) **recall** trends in the properties of the fourth row transition elements.

14.3.b(1) **list** some of the properties of chromium, manganese, iron, nickel, copper and zinc.

b(2) **explain** the separation of the above ions in terms of the stability of these cation or anion complexes.

b(3) **recall** the factors which promote the corrosion of iron.

b(4) **explain** the corrosion of iron using redox equations.

b(5) **perform** an experiment to illustrate the factors involved in the corrosion of iron.

b(6) **solve** problems related to metal corrosion.

b(7) **define** a sacrificial anode.

UNIT 15

SOME SIXTH-ROW AND SEVENTH-ROW ELEMENTS

At the conclusion of this unit, the student should be able to:

15.1.a(1) **locate** the lanthanides (rare earths) on the periodic table.

a(2) **identify** the electron energy sub-level filled in the lanthanide series.

a(3) **recognize** the similar chemical properties of the lanthanides.

a(4) **explain** why the lanthanides have a uniform 3^+ oxidation number.

15.1.b(1) **locate** the actinides on the periodic table.

b(2) **identify** the electron energy sub-level filled in the series.

b(3) **recognize** that the seventh-row elements are radioactive.

15.1.c(1) **define:**

(i) radioactivity
(ii) radioisotope
(iii) half-life

c(2) **state** methods used to detect radioactivity.

c(3) **solve** problems relating to half-life.

c(4) **relate** the stability of a nucleus to its half-life.

c(5) **relate** the stability of a nucleus to its neutron to proton ratio.

c(6) **describe** the uses of radioisotopes.

15.1.d(1) **write** the symbol for nuclear particles (e.g. neutron (^1_0n))

d(2) **balance** and/or **complete** nuclear equations.

d(3) **identify** the types of radioactivity (e.g. -decay).

15.1.e(1) **solve** problems using Einstein's mass-energy relationship ($E = mc^2$).

e(2) **define:**

(i) nuclear binding energy.
(ii) nuclear binding energy per particle.
(iii) nuclear fusion.
(iv) nuclear fission.
(v) nuclear chain reaction,
(vi) moderator

e(3) **solve** problems involving nuclear binding energy.

e(4) **recognize** that the nuclear binding energy per particle is a measure of the stability of the nucleus.

e(5) **explain** nuclear fusion and fission in terms of changes in binding energy per particle.

e(6) **describe** how a nuclear fission reaction is controlled.

e(7) **discuss** nuclear reactions as sources of energy.

e(8) **discuss** the feasibility of alternate sources of energy.

e(9) **discuss** the health hazards of radioactive materials.

e(10) **discuss** the role of society and government in the development of nuclear policy.

(b) S-17E MULTIPLE CHOICE INSTRUMENTS

13A.1.

1) X, Y and Z are experimentally measured quantities. In a particular experiment the measured values are

$$X = 100 \text{ m} \pm 10\%$$
$$Y = 50 \text{ m} \pm 10\%$$
$$Z = 5 \text{ s} \pm 10\%$$

$$P = \frac{X - Y}{Z} = \frac{100 - 50}{5} = 10 \text{ m/s}$$

The uncertainty in the calculated value of P is

A) $\pm$ 10%
B) $\pm$ 20%
C) $\pm$ 30%
*D) $\pm$ 40%

2) 1 N is equivalent to

*A) 1 $kg \cdot m \cdot s^{-2}$
B) 1 $kg \cdot m^{-1} \cdot s^{-2}$)
C) 1 $kg \cdot s^{2} \cdot m^{-1}$
D) 1 $s^{2} \cdot kg^{-1} \cdot m^{-1}$

3) One cubic metre of water is equal in volume to

A) 10 dm^3 (cubic decimetre)
B) 100 dm^3 (cubic decimetre)
*C) 1000 dm^3 (cubic decimetre)
D) 10000 dm^3 (cubic decimetre)

13A.1.

4) 15.2 ± 0.2 g of element A has a measured volume of 1.12 ± 0.03 mL. Element A has a density of

A) 13.6 ± 0.2 g/mL
B) 13.6 ± 0.3 g/mL
*C) 13.6 ± 0.5 g/mL
D) 13.6 ± 0.7 g/mL

5) Breaking a compound down into its elements is known as

*A) decomposition
B) fission
C) synthesis
D) transmutation

6) Rutherford is best known for the

*A) nuclear model of the atom
B) the charge/mass ratio of an electron
C) development of quantum theory
D) discovery of X-rays

7) An atom which contains 19 neutrons has a mass number of 35. How many **protons** does this atom contain?

*A) 16
B) 19
C) 35
D) 54

8) The symbol $^{238}_{92}U$ indicates that each uranium atom contains

A) 238 protons and 92 electrons
*B) 92 protons and 146 neutrons
C) 92 protons and 238 neutrons
D) 146 protons and 92 neutrons

13A.1.

9) An element has two isotopes, X and Y, of atomic masses 52.0 and 53.0 respectively. The relative atomic mass of the element is 52.2. The percentage abundance of the two isotopes in nature is

	X	Y
A)	20%,	80%
B)	25%,	75%
*C)	80%,	20%
D)	75%,	25%

10) An ion of the element Al (aluminum) carries a charge of 3+. Which choice in the following table best describes this ion?

	ATOMIC No.	MASS No.	No.OF PROTONS	No.OF NEUTRONS	No.OF ELECTRONS
A)	10	13	13	10	13
B)	13	13	10	14	13
*C)	13	27	13	14	10
D)	13	27	13	14	13

11) In chemical reactions negative ions are formed from neutral atoms by

A) loss of electrons
B) loss of protons
C) gain of protons
*D) gain of electrons

12) The smallest particle of any pure substance able to exist physically and still retain the original composition of that pure substance is called a(n)

A) atom
B) ion
*C) molecule
D) electron

13A.1.

13) What is the name of the process which involves the physical change of a substance directly from solid to gas?

A) evaporation
B) liquification
*C) sublimation
D) condensation

14) One Pascal is equal to

A) one square metre per newton
B) one metre per newton
*C) one newton per square metre
D) one newton per metre

15) Charles' Law states that at constant pressure and mass

A) the volume of a gas will be inversely proportional to the Kelvin temperature
*B) the volume of a gas will be directly proportional to the Kelvin temperature
C) the volume of a gas is independent of the Kelvin temperature
D) the volume of a gas will be directly proportional to pressure and Kelvin temperature

16) Consider an ideal gas with volume of V at a particular temperature and pressure. If the Kelvin temperature is doubled and the pressure is quadrupled, the new volume of the gas, expressed in litres, is

*A) $\frac{V}{2}$
B) V
C) 2V
D) 4V

13A.1.

17) A gas has a volume of 16 mL at 364 K. If the mass and pressure remain unchanged, the volume of the gas at 273 K will be

A) 3.0 mL
*B) 12 mL
C) 21 mL
D) 32 mL

18) A 3.0 g sample of a gas, collected at 363 K and 101.3 kPa, occupies 2.0 L. The density of this gas at STP is

A) 1.0 g/L
*B) 2.0 g/L
C) 3.0 g/L
D) 0.5 g/L

19) A gas has a volume of 45 L at 733 K and 1.5×10^2 kPa. The volume of the gas, at STP, is

A) 15 L
*B) 25 L
C) 81 L
D) 1.4×10^2 L

20) A fixed mass of gas has a volume of 4.0×10^2 L and a pressure of 1.0×10^2 kPa. If the pressure is reduced to 50 kPa at constant temperature, the volume of the gas is

*A) 8.0×10^2 L
B) 2.0 L
C) 1.2×10^2 L
D) 4.0×10^2 L

13A.1.

21) At 101 kPa of pressure, 20.0 L of oxygen gas is placed in a cylinder having a movable piston. What pressure must be applied to the piston to reduce the volume to 2.00 L if the temperature is held constant?

A) 1.01 kPa
B) 9.09×10^2 kPa
*C) 1.01×10^3 kPa
D) 2.01×10^3 kPa

22) At 300 K and 404 kPa pressure, a sample of oxygen gas has a mass of 4.0 g and occupies a volume of 3.0 L. The density of the gas at 200 K and 202 kPa will be

*A) 1.0 g/L
B) 2.0 g/L
C) 3.0 g/L
D) 4.0 g/L

23) Water reacts with the oxides of non-metals to form

A) basic solutions
B) anhydrides
C) neutral solutions
*D) acidic solutions

24) 2.00 ± 0.05 mol of element A has a mass of 92.6 ± 0.5 g. The molar mass of A is

A) 46.3 ± 0.55 g
B) 46.3 ± 1.1 g
*C) 46.3 ± 1.4 g
D) 46.3 ± 2.5 g

25) The number of **molecules** of carbon dioxide in 11.0 g of CO_2 (carbon dioxide) is

*A) 1.50×10^{23}
B) 2.24×10^{23}
C) 3.01×10^{23}
D) 6.02×10^{23}

26) How many moles of water are there in 500 g of water (H_2O)?

*A) 27.8 mol
B) 29.4 mol
C) 55.6 mol
D) 9.00×10^3 mol

27) How many moles of $Mg(NO_3)_2$ (magnesium nitrate) are present in 40 g of $Mg(NO_3)_2(s)$?

*A) 0.27
B) 0.37
C) 0.54
D) 3.7

28) A sample of $O_2(g)$ (oxygen) at STP contains 2.4 mol of molecules. The sample would have a volume of

A) 9.0 L
B) 25 L
*C) 54 L
D) 90 L

13A.1.

29) How many moles of HNO_3 (nitric acid) are contained in 157.5 g of nitric acid?

A) 0.40
B) 2.0
*C) 2.5
D) 4.0

30) The percentage (by mass) chlorine in $HClO_3$ (chloric acid) is

A) 24%
B) 36%
*C) 42%
D) 58%

31) What is the simplest (empirical) formula of a compound of carbon and hydrogen 10.0 g of which contains 7.50 g carbon?

A) C_2H
*B) CH_4
C) CH_2
D) C_2H_2

32) The number of atoms of carbon in one mole of C_2H_6 (ethane) is

A) 3.0×10^{23}
B) 6.0×10^{23}
*C) 1.2×10^{24}
D) 1.8×10^{24}

13A.1.

33) If 1 L of a gas at STP has a mass of 2.86 g the molar mass of the gas is

A) 2.86 g
B) 22.4 g
*C) 64.1 g
D) 286 g

34) The volume occupied by 10 g of fluorine gas (F_2) at 300 K and 82 kPa pressure is

*A) 8.0 L
B) 10 L
C) 16 L
D) 22 L

35) Compressed oxygen gas is sold in steel cylinders of capacity 40 L. The pressure in the cylinders is 1.9×10^4 kPa and the temperature is $27^{\circ}C$. The number of moles of oxygen gas in the cylinder is

A) 3.0×10^{-2}
B) 3.0×10^{-1}
C) 3.0×10^{1}
*D) 3.0×10^{2}

36) The names of the compound ions

$$HCO_3^-,\ NO_3^-,\ OH^-,\ CrO_4^{2-}$$

are

A) carbonate, nitrate, hydride, dichromate
B) carbonate, nitrite, hydroxide, cyanate
*C) hydrogen carbonate, nitrate, hydroxide, chromate
D) hydrogen carbonate, nitrite, hydride, cyanate

37) The number of moles of oxygen gas ($O_2(g)$) produced when 4 mol of hydrogen gas ($H_2(g)$) is produced by the electrolysis of water is

A) 6 mol
*B) 2 mol
C) 8 mol
D) 4 mol

38) Consider the following balanced chemical equation:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g) + \text{energy}$$

What volume of O_2 (oxygen gas) will combine with 1.0 L of H_2 (hydrogen gas) to form water, both gases being measured at the same temperature and pressure?

*A) 0.50 L
B) 2.0 L
C) 8.0 L
D) 16 L

39) 37 g of $Ca(OH)_2$ (calcium hydroxide) reacts as indicated by the equation:

$$Ca(OH)_2 + CO_2 \rightarrow CaCO_3 + H_2O\ (g)$$

The volume of water vapour that would be produced if the reaction had occurred at 4.0×10^2 K and 2.0×10^2 kPa pressure is

A) 5.6 L
*B) 8.2 L
C) 16 L
D) 22 L

13A.1.

40) What mass of $CaCO_3$ (calcium carbonate) is required to produce 84 g of CaO (calcium oxide)?

$$CaCO_3 \rightarrow CaO + CO_2$$

A) 42 g
B) 67 g
C) 100 g
*D) 150 g

41) Water can be decomposed by electrolysis to yield $H_2(g)$ (hydrogen gas) and $O_2(g)$ (oxygen gas).

$$2H_2O(\ell) \rightarrow 2H_2(g) + O_2(g)$$

If 56 mL of O_2 are produced (measured at STP), the number of moles of water that are decomposed is

A) 9×10^{-2} mol
*B) 5×10^{-3} mol
C) 2.5×10^{-3} mol
D) 5×10^{-4} mol

42) $KClO_3$ (potassium chlorate) decomposes according to the equation:

$$2KClO_3 \rightarrow 2KCl + 3O_2$$

If 12.25 g of $KClO_3$ is totally decomposed, what mass of O_2 is produced?

A) 1.6 g
B) 3.2 g
*C) 4.8 g
D) 9.6 g

13A.1.

43) What mass of $CaCO_3$ (calcium carbonate) would be required to react completely with HBr (hydrogen bromide) to form 50 g of $CaBr_2$ (calcium bromide)?

$$CaCO_3 + 2HBr \rightarrow CaBr_2 + CO_2 + H_2O$$

*A) 25 g
B) 50 g
C) 1.0×10^2 g
D) 2.0×10^2 g

44) Consider the equation:

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

How many litres of $H_2(g)$ (hydrogen gas) measured at 101 kPa pressure and $0^{\circ}C$ are required to react completely with 11.2 L of $N_2(g)$ (nitrogen gas) at STP?

A) 16.8 L
B) 22.4 L
*C) 33.6 L
D) 67.2 L

45) Which process will purify water by separating from it all the dissolved solids?

A) boiling
B) filtration
C) sedimentation
*D) distillation

13A.1.

46) 20 g of NaOH (sodium hydroxide) is dissolved in 100 mL of distilled water and the solution is made up to 500 mL with water. The resultant concentration, in mol/L, is

*A) 1.0
B) 2.0
C) 5.0
D) 10

47) The concentration of a solution containing 100 g of NaOH in 2.5 L of solution is

*A) 1.0 mol/L
B) 2.5 mol/L
C) 40 mol/L
D) 4.0 mol/L

48) The volume of 6.0 mol/L HCl (hydrochloric acid) which must be diluted with water to prepare 1.0 L of 1.2 mol/L HCl is

A) 7.2 mL
B) 20 mL
*C) 2.0×10^2 mL
D) 7.2×10^2 mL

49) What mass of HCl (hydrochloric acid) is required to make 1.00 L of a 0.100 mol/L solution?

A) 0.100 g
*B) 3.65 g
C) 18.2 g
D) 36.5 g

13A.1.

50) During the laboratory preparation of H_2 (hydrogen) according to the equation:

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2$$

*A) the zinc atoms are oxidized
B) the hydrogen ions are oxidized
C) the chloride ions are reduced
D) the zinc atoms are reduced

51) In the reaction:

$$MnO_4^-(aq) + Fe^{2+}(aq) + H^+(aq) \rightarrow Mn^{2+}(aq) + Fe^{3+}(aq) + {}_2H\,O(\ell)$$

the reactant species that was oxidized is

A) MnO_4^-

B) Mn^{7+}

*C) Fe^{2+}

D) H^+

52) A common characteristic of solutions of **both** strong acids and strong bases is that they

A) turn litmus red
B) react with magnesium to produce hydrogen
*C) conduct electricity
D) contain a surplus of OH^-

53) The oxide compounds of the **group 2** metals, when dissolved in water, generally produce

*A) bases
B) alkaline earths
C) acids
D) salts

13A.1.

54) The periodic table is a classification of the elements in order of increasing number of

A) neutrons
B) positrons
*C) protons
D) isotopes

55) Element 19 is chemically more active than element 11 because element 19 has

A) a greater number of valence electrons
B) a smaller number of valence electrons
*C) valence electrons farther from the nucleus
D) valence electrons closer to the nucleus

56) Elements Q, R, X and T occur in the same period of the Periodic Table and have two, three, four and seven valence electrons, respectively. Which element is most likely to exist in the form of diatomic molecules?

A) Q
B) R
C) X
*D) T

57) At STP, CO (carbon monoxide), N_2 (nitrogen) and Ne (neon) are all

A) solids
B) liquids
*C) gases
D) solutions

13A.1.

58) The energy required to remove an electron from an atom under specified conditions is called

A) the electronic energy
B) the atomic energy
*C) the ionization energy
D) the electronegativity

59) An example of a pure substance in which the only forces between atoms are van der Waals forces is

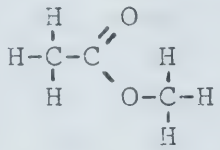
A) CH_4 (methane)
B) Na (sodium)
*C) He (helium)
D) H_2 (hydrogen)

60) The bonding in crystalline KBr (potassium bromide) is called

A) metallic
B) covalent
C) molecular
*D) ionic

13A.1.

61) The ester represented by the structural formula



can be synthesized from

A) acetic acid and ethanol
B) formic acid and ethanol
C) formic acid and methanol
*D) acetic acid and methanol

62) The density of O_2 (oxygen) gas at STP is 1.43 g/L. What would its density be at 373 K and 98.1 kPa?

*A) 1.02 g/L
B) 1.46 g/L
C) 1.89 g/L
D) 2.02 g/L

63) The number of **molecules** contained in 22.4 L of oxygen gas at 273 K and 202.6 kPa pressure would be

A) 3.01×10^{23}
B) 6.02×10^{23}
*C) 1.20×10^{24}
D) 2.4×10^{24}

64) The solubility of potassium chlorate ($KClO_3$) in water at room temperature is 10.0 g per **100 g of water.** The percent, by mass, of solute in **the solution** is

*A) 9.09%
B) 10.0%
C) 11.1%
D) 14.3%

13A.1.

65) A sample of methane gas (CH_4) at STP, contains 1.7 mol of molecules. The volume of the sample is

A) 0.076 L
B) 13 L
C) 24 L
*D) 38 L

66) The molar mass of calcium phosphate ($Ca_3(PO_4)_2$) is

A) 150 g
B) 200 g
*C) 310 g
D) 470 g

67) A sample of hydrogen gas contains a small amount of the isotope deuterium, D. In a mass spectrometer the ion HD^+ is formed. The relative atomic mass of D is 2. The HD^+ ion contains:

A) 1 proton, 1 neutron, 1 electron
B) 2 protons, 1 neutron, 2 electrons
C) 2 protons, 2 neutrons, 1 electron
*D) 2 protons, 1 neutron, 1 electron

68) Which of the following are isotopes of the same element

I Z = 10, A = 20
II Z = 11, A = 21
III Z = 10, A = 21
IV Z = 12, A = 20

A) I and IV
B) II and III
C) I and II
*D) I and III

13A.1.

69) An atom of the "element" whose atomic number is 120 forms a 2+ ion. The number of protons and electrons in the ion is

A) 118 p and 120 e
*B) 120 p and 118 e
C) 120 p and 120 e
D) 122 p and 120 e

70) A sample of a compound contains 14.4 g of calcium, 4.32 g of carbon and 17.3 g of oxygen. The simplest (empirical) formula for the compound is

A) CaCO
B) $CaCO_2$
*C) $CaCO_3$
D) Ca_3CO_3

71) An element consists of two isotopes of relative atomic masses 41.0 and 44.0. The relative abundance of the isotopes are 90% and 10% respectively. The relative atomic mass of the element is

A) 41.0
*B) 41.3
C) 42.5
D) 43.7

72) For which of the following pairs of anions do both names end with the suffix "ate"?

A) F^- ClO^-
B) ClO_3^- ClO_2^-
*C) CO_3^{2-} SO_4^{2-}
D) SO_3^{2-} MnO_4^-

13A.1.

73) A sample of freon gas (CCl_2F_2) has a mass of 30 g. The number of moles of molecules of freon in the 30 g sample is

*A) 0.25
B) 0.45
C) 2.2
D) 4.0

74) What is the concentration, in mol/L, of a solution of sulfuric acid (H_2SO_4) that contains 15 g of pure H_2SO_4 in 2.0×10^2 mL of solution?

A) 0.15
B) 0.63
*C) 0.77
D) 0.83

75) Atoms of copper usually form ions with charges of either

A) +2 or +3
*B) +1 or +2
C) -1 or +1
D) -1 or -2

76) One atom of the isotope $^{35}_{17}Cl$ has a nucleus containing

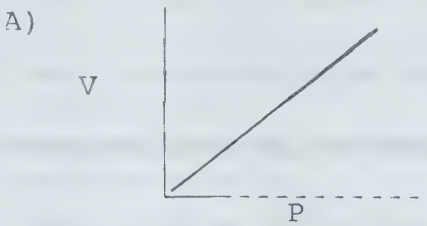
A) 17 protons and 35 neutrons
B) 35 protons and 17 neutrons
*C) 17 protons and 18 neutrons
D) 18 protons and 17 neutrons

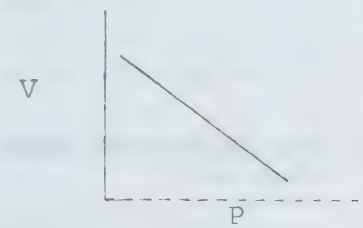
13A.1.

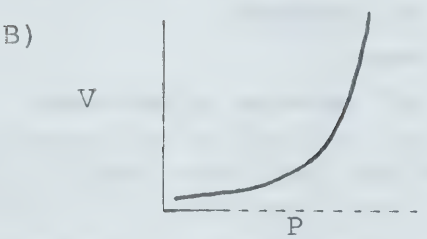
77) The formula of the compound usually formed by the reaction of magnesium and sulphur is

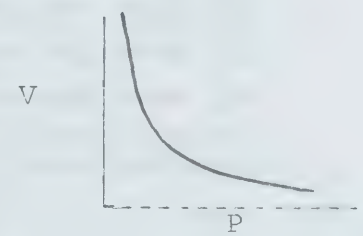
*A) MgS
B) Mg_2S
C) MgS_2
D) Mg_6S_2

78) Which graph best illustrates Boyle's Law?

A)


C)


B)


*D)


79) 20 mL of HCl (concentration = 6.0 mol/L) is added to sufficient water to make 100 mL of solution. What is the HCl concentration in the resulting solution?

A) 0.060 mol/L
B) 0.60 mol/L
*C) 1.2 mol/L
D) 1.5 mol/L

13A.1.

80) 2.17 g of mercury(II) oxide (HgO, molar mass = 217 g/mol) is heated to produce mercury and oxygen:

$$2HgO(s) \rightarrow 2Hg(\ell) + O_2(g)$$

How many moles of $O_2(g)$ will be produced in the reaction?

*A) 5.00×10^{-3} mol
B) 1.00×10^{-2} mol
C) 2.00×10^{-2} mol
D) 5.00×10^{1} mol

81) The halogens have similar chemical properties because they all

A) are coloured gases
B) have the same number of protons
C) have the same atomic mass
*D) have the same number of valence electrons

82) The correct name of the compound $Cu(ClO_3)_2$ is

A) copper(II) perchlorate
B) copper(I) chlorate
*C) copper(II) chlorate
D) copper(I) hypochlorite

83) Given that the relative atomic masses of hydrogen and oxygen are 1.008 and 15.9994 the molar mass of water (H_2O), to the appropriate number of significant figures, is

A) 18.01 g/mol
B) 18.02 g/mol
C) 18.0154 g/mol
*D) 18.015 g/mol

13A.1.

84) When examining the Periodic Table, one would expect to find an increasing pattern in electropositive behaviour as the observer moves

*A) down and to the left
B) down and to the right
C) up and to the left
D) up and to the right

85) In the reaction between potassium and sulphur

A) each potassium atom donates two electrons to a sulphur atom
*B) each potassium atom donates one electron to a sulphur atom
C) each sulphur atom donates one electron to a potassium atom
D) each sulphur atom donates two electron to a potassium atom

86) The balanced equation for the reaction between aqueous $Pb(NO_3)_2$ (lead nitrate) and aqueous NaI (sodium iodide) is

A) $Pb(NO_3)_2 + NaI \rightarrow PbI + NaNO_3$
B) $Pb(NO_3)_2 + 2NaI \rightarrow PbI_2 + Na_2NO_3$
C) $Pb(NO_3)_2 + 2NaI \rightarrow 2PbI + 2NaNO_3$
*D) $Pb(NO_3)_2 + 2NaI \rightarrow PbI_2 + 2NaNO_3$

87) What is the solubility of a salt in grams of solute per 100 g of solvent if the salt forms a saturated solution containing 6.00 g of salt in 50.0 g of solution at 298 K?

A) 6.00
B) 12.0
*C) 13.6
D) 25.0

13A.1.

88) A concentration of I_2 (iodine) dissolved in alcohol is 0.20 mol/L; the density of the solution is 0.85 g/mL. The relative molecular mass of I_2 is 254. The percentage by mass of I_2, iodine, in this solution is

A) 2.4%
B) 3.0%
C) 5.6%
*D) 6.0%

89) What mass of KOH(s) (potassium hydroxide) is needed to prepare 2.0 L of a 2.0 mol/L KOH solution?

A) 1.1×10^2 g
B) 2.0×10^2 g
*C) 2.2×10^2 g
D) 4.0×10^2 g

90) The solubility of NaCl (sodium chloride) in water at 333 K is 38.0 g per 100 g of water. The percent by mass of solute in the solution is

A) 26.3%
*B) 27.5%
C) 36.3%
D) 38.0%

91) A solution of NaOH (sodium hydroxide) is 20% solute by mass and has a density of 1.1 g/mL. The concentration NaOH of the solution is

A) 1.8 mol/L
B) 3.6 mol/L
C) 4.2 mol/L
*D) 5.5 mol/L

13A.1.

92) A solution of I_2 (iodine) in alcohol has a density of 0.90 g/mL. The mass of 50 mL of the solution is

A) 41 g
*B) 45 g
C) 51 g
D) 56 g

93) A salt forms a saturated solution which contains 15.0% solute by mass. What is the solubility of this salt in g of solute per 100 g of solvent?

A) 11.3
B) 13.0
C) 15.0
*D) 17.6

94) The balanced net ionic equation for the reaction of lead(II) nitrate and iron(II) chloride is

A) $Pb^{2+}(aq) + 2Cl^{-}(aq) = Pb(s) + Cl_2(g)$

B) $Fe^{2+}(aq) + 2NO_3^{-}(aq) = Fe(NO_3)_2(s)$

*C) $Pb^{2+}(aq) + 2Cl^{-}(aq) = PbCl_2(s)$

D) $2Fe^{2+}(aq) + Pb^{2+}(aq) = Pb(s) + 2Fe^{3+}(aq)$

13A.2.

1) Which one of the following reactions involves the **smallest** amount of energy per mole of product?

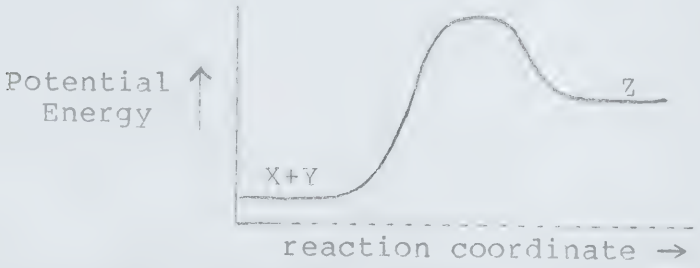
A) $^{231}_{91}Pa \rightarrow {}^{4}_{2}He(\ell) + {}^{227}_{89}Ac$ + energy

B) $N_2(g) \rightarrow N_2(\ell)$ + energy

*C) $Ag^+(aq) + Cl^-(g) \rightarrow AgCl(s)$ + energy

D) $2H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$ + energy

2) The following diagram shows which fact about the reaction X + Y → Z?

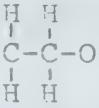


A) It is rapid.
*B) It is endothermic.
C) It is exothermic.
D) It is spontaneous.

13A.2.

3) Given the following bond energies:

C - H	416 kJ/mol
C - O	361 kJ/mol
O - H	466 kJ/mol
C - C	349 kJ/mol

The bonding energy for  (ethyl alcohol) is

A) 1.59×10^3 kJ/mol

B) 2.01×10^3 kJ/mol

C) 2.84×10^3 kJ/mol

*D) 3.26×10^3 kJ/mol

4) "When a particular reaction can be expressed as the sum of other reactions, the heat of the reaction is the algebraic sum of the heats of these reactions." This statement is referred to as

A) Nobel's Law

*B) Hess' Law

C) Boyle's Law

D) Priestley's Law

5) 0.25 mol of element X was burned in oxygen to produce 0.25 mol of the corresponding oxide. The temperature of 200 g of surrounding water rose 15°C. The specific heat capacity of water is 4.2 J/(g·°C). The $\triangle H_f$ of the oxide of X is

A) -1.26 kJ/mol

B) -3.15 kJ/mol

C) -12.6 kJ/mol

*D) -50.4 kJ/mol

13A.2.

6) $C_2H_6(g) \rightarrow C_2H_2(g) + 2H_2(g) \quad \triangle H = +\ 3.2 \times 10^2$ kJ/mol (C_2H_6)

An alternate way to write this equation is

A) $C_2H_6(g) \rightarrow C_2H_2(g) + 2H_2(g) + 3.2 \times 10^2$ kJ

*B) $C_2H_6(g) + 3.2 \times 10^2$ kJ $\rightarrow C_2H_2(g) + 2H_2(g)$

C) $2H_2(g) + C_2H_2(g) + 3.2 \times 10^2$ kJ $\rightarrow C_2H_6(g)$

D) $C_2H_6(g) - 3.2 \times 10^2$ kJ $\rightarrow C_2H_2(g) + 2H_2(g)$

7) Consider the following equation for the combustion of H_2(hydrogen):

$$H_2(g) + \tfrac{1}{2}O_2(g) \rightarrow H_2O(g) + 243 \text{ kJ}$$

In order to produce 972 kJ of heat, how many grams of H_2 must burn?

A) 0.250 g
B) 2.00 g
*C) 8.00 g
D) 16.00 g

8) 7.28 kJ of heat is released when 40.0 g of Br_2 (bromine) vapour at 332 K condense to a liquid. What is the molar heat of vaporization of Br_2?

A) 3.64 kJ/mol
B) 7.28 kJ/mol
C) 14.6 kJ/mol
*D) 29.1 kJ/mol

13A.2.

9) The $\triangle H$ for a reaction is 77 kJ/mol and the $\triangle S$ for the reaction is 22 kJ/(mol·K). The temperature at which the reaction becomes spontaneous is

*A) 3.5 K
B) 3.5×10^3 K
C) 55 K
D) 1.7×10^3 K

10) The reaction which represents the **greatest** increase in entropy is

*A) $X_2(s) + Y_2(g) \rightarrow 2XY(g)$
B) $M(s) + O(s) \rightarrow MO(s)$
C) $X(g) \rightarrow X(\ell)$
D) $X_2(g) + Y_2(g) \rightarrow 2XY(g)$

11) Which of the following reactions are exothermic?

I) $H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(g) \qquad \triangle H = -243$ kJ/mol

II) $\frac{1}{2} N_2(g) + \frac{1}{2} O_2(g) + 90.7$ kJ $\rightarrow NO(g)$

III) $C(s) + \frac{1}{2} O_2(g) \rightarrow CO(g) + 111$ kJ

IV) $\frac{1}{2} H_2(g) + \frac{1}{2} I_2(g) \rightarrow HI(g) \qquad \triangle H = +26$ kJ/mol

A) I and II only
B) I and IV only
*C) I and III only
D) II and IV only

13A.2.

12) Given the following thermochemical data:

1) $C_2H_2(g) + \frac{5}{2}O_2(g) \rightarrow 2CO_2(g) + H_2O(\ell)$ $\triangle H_1 = -1.30 \times 10^3$ kJ

2) $C_2H_6(g) + \frac{7}{2}O_2(g) \rightarrow 2CO_2(g) + 3H_2O(\ell)$ $\triangle H_2 = -1.56 \times 10^3$ kJ

3) $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(\ell)$ $\triangle H_3 = -2.86 \times 10^2$ kJ

What is $\triangle H$ for the reaction:

$$C_2H_2(g) + 2H_2(g) \rightarrow C_2H_6(g) \quad \triangle H = ?$$

A) -2.60×10^1 kJ
*B) -3.12×10^2 kJ
C) -5.72×10^2 kJ
D) -5.46×10^3 kJ

13) $NaCl(s) + H_2O(\ell) \rightarrow NaCl(aq)$ $\triangle H = +1.67$ kJ/mol of NaCl (at 298 K)

If sufficient NaCl is dissolved in water (at 298 K) to form 1.00 L of 5.00 mol/L NaCl solution, the change in enthalpy of the system will be

A) 1.67 kJ
B) 3.36 kJ
*C) 8.35 kJ
D) 21.0 kJ

14) The increase in vapour pressure of a liquid caused by an increase in temperature is best explained by which one of the following statements?

A) More molecules can exist in the gaseous state at a higher temperature but the average kinetic energy is not changed.
*B) More molecules can exist in the gaseous state at a higher temperature and the average kinetic energy is higher.
C) The same number of molecules can exist in the gaseous state at a higher temperature but the average kinetic energy is lower.
D) The increase in temperature causes an increase in mass of each molecule and therefore a higher pressure.

13A.2.

15) Consider the enthalpy changes for the following reactions:

1) $C(s) + 2H_2(g) \rightarrow CH_4(g)$ $\quad \triangle H_1 = -75.60$ kJ

2) $C(s) + 2H_2(g) \rightarrow C(g) + 4H(g)$ $\quad \triangle H_2 = +380.0$ kJ

What is the C - H bond energy?

A) -18.9 kJ/mol
*B) 113.9 kJ/mol
C) 304.4 kJ/mol
D) 455.6 kJ/mol

16) Consider the following four equations:

1) $C_6H_6(\ell) + \frac{15}{2}O_2(g) \rightarrow 6CO_2(g) + 3H_2O(\ell)$ $\quad \triangle H_1$

2) $6C(s) + 3H_2(g) \rightarrow C_6H_6(\ell)$ $\quad \triangle H_2$

3) $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(\ell)$ $\quad \triangle H_3$

4) $C(s) + O_2(g) \rightarrow CO_2(g)$ $\quad \triangle H_4$

The enthalpy change for reaction (1), $\triangle H_1$, can be obtained by algebraically combining the $\triangle H$'s for the other reactions in which of the following ways?

A) $6(\triangle H_4) + 3(\triangle H_3) + \triangle H_2$
B) $6(\triangle H_4) - 3(\triangle H_3) + \triangle H_2$
C) $-6(\triangle H_4) - 3(\triangle H_3) + \triangle H_2$
*D) $6(\triangle H_4) + 3(\triangle H_3) - \triangle H_2$

17) The spontaneous evaporation of methanol (CH_3OH) is evidence of

A) a natural tendency towards minimum potential energy
*B) a natural tendency towards increasing randomness
C) the existence of a state of equilibrium
D) a high activation energy

13A.2.

18) The molar heat of vaporization of water is 42 kJ/mol. How much energy is released by the condensation of 3.0 g of water?

A) 0.88 kJ
*B) 7.0 kJ
C) 130 kJ
D) 250 kJ

19) The heat of formation of sodium chlorate ($NaClO_3(s)$) is -53 kJ/mol. The balanced equation which expresses this information is

A) $Na(s) + Cl_2(g) + O_2(g) \rightarrow NaClO_3(g) + 53\ kJ$

*B) $Na(s) + \frac{1}{2}Cl_2(g) + \frac{3}{2}O_2(g) \rightarrow NaClO_3(s) + 53\ kJ$

C) $Na(s) + \frac{1}{2}Cl_2(g) + \frac{3}{2}O_2(g) + 53\ kJ \rightarrow NaClO_3(s)$

D) $Na(s) + Cl_2(g) + 3O_2(g) \rightarrow NaClO_3(g) + 53\ kJ$

20) The heat of formation of magnesium oxide (MgO) from the elements is -603 kJ/mol. How much heat is required to decompose 24.0 g of MgO?

A) 181 kJ
*B) 362 kJ
C) 723 kJ
D) 1.0×10^3 kJ

21) The heat of combustion of methane (CH_4) is -891 kJ/mol. What mass of methane must be burned to produce enough heat to cook 50.0 g of potatoes which require 530 J/g for cooking to be complete?

*A) 0.480 g
B) 0.540 g
C) 33.6 g
D) 480 g

13A.2.

22) In an exothermic reaction

A) potential energy is transformed into bonding energy
B) heat is transformed into potential energy
*C) potential energy is transformed into kinetic energy
D) kinetic energy is transformed into bonding energy

23) The products of combustion of H_2S (hydrogen sulfide) in oxygen are H_2O (water) and SO_2 (sulfur dioxide). Given the following data, what is the heat of combustion of H_2S?

$$H_2(g) + S(s) \rightarrow H_2S(g) \qquad \triangle H = -40.2 \text{ kJ/mol}$$
$$H_2(g) + \tfrac{1}{2}O_2(g) \rightarrow H_2O(g) \qquad \triangle H = -242 \text{ kJ/mol}$$
$$S(s) + O_2(g) \rightarrow SO_2(g) \qquad \triangle H = -297 \text{ kJ/mol}$$

A) -15 kJ/mol
B) -95 kJ/mol
*C) -499 kJ/mol
D) -579 kJ/mol

24) Which of the following combinations of enthalpy and entropy changes would **always** give a spontaneous reaction?

	enthalpy	entropy
A)	increased	increased
B)	increased	decreased
C)	decreased	decreased
*D)	decreased	increased

13A.2.

25) Consider the following exothermic reactions:

I $2H(g) \rightarrow H_2(g)$ + heat

II $H_2(g) \rightarrow H_2(s)$ + heat

III $H_2(\ell) \rightarrow H_2(s)$ + heat

IV $^{3}_{1}H + {}^{2}_{1}H \rightarrow {}^{4}_{2}He + {}^{1}_{0}n$ + heat

The reactions in order of increasing magnitude of their expected $\triangle H$'s (i.e., lowest $\triangle H$ to highest $\triangle H$) is

A) II < III < IV < I
*B) III < II < I < IV
C) IV < I < II < III
D) I < III < II < IV

26) Consider the following data:

$C(s) + \frac{1}{2}O_2(g) \rightarrow CO(g)$ $\quad \triangle H = -111$ kJ

$C(s) + O_2(g) \rightarrow CO_2(g)$ $\quad \triangle H = -395$ kJ

The enthalpy change per mole of CO_2 for the following reaction is

$$CO(g) + \frac{1}{2}O_2(g) \rightarrow CO_2(g)$$

A) -506 kJ
*B) -284 kJ
C) +284 kJ
D) +506 kJ

27) For the reaction between $H_2O(g)$ and $SO_2(g)$ (at 298 K) $\triangle H$ is +143.6 kJ/mol and $\triangle S$ is 0.0501 kJ/(mol·deg).

$$H_2O(g) + SO_2(g) \rightarrow H_2(g) + SO_3(g)$$

$\triangle G$, the free energy change, for the reaction is

A) -158.5 kJ/mol
B) -128.7 kJ/mol
*C) 128.7 kJ/mol
D) 143.6 kJ/mol

28) The spontaneous evaporation of CH_3OH (methanol) is accompanied by absorption of 1.18 kJ/g of methanol. What is $\triangle H$ for the reaction:

$$CH_3OH(\ell) \rightarrow CH_3OH(g)$$

A) 0.0369 kJ/mol
B) 1.18 kJ/mol
C) 32.0 kJ/mol
*D) 37.8 kJ/mol

29) Consider the following exothermic reaction:

$$^{3}_{1}H + ^{2}_{1}H \rightarrow ^{4}_{2}He + ^{1}_{0}n + \text{heat}$$

This process would most likely occur at a temperature of

A) 13 K
B) 6.0×10^2 K
C) 6.0×10^3 K
*D) 6.0×10^7 K

13A.2.

30) The heat released during the combustion of $CH_4(g)$ (methane) to produce CO_2(carbon dioxide) gas and water vapour is 882 kJ/mol of $CH_4(g)$. Which of the following is the correct chemical equation for the reaction?

A) $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O\ (g)$ $\quad \triangle H = +882$ kJ

B) $CH_4(g) + 2O_2(g) + 882\ kJ \rightarrow 2H_2O\ (g) + CO_2(g)$

*C) $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O\ (g) + 882$ kJ

D) $CO_2(g) + 2H_2O\ (g) \rightarrow 2O_2(g) + CH_4(g)$ $\quad \triangle H = -882$ kJ

31) 2.00 mol of an element burned in a calorimeter raises the temperature of 200 g of water 10.0°C. (Specific heat of water = 4.18 kJ/(kg·°C) The molar heat of combustion of the element is

A) 0.239 kJ/mol

*B) 4.18 kJ/mol

C) 8.36 kJ/mol

D) 16.7 kJ/mol

32) When the temperature of a system is low, the reaction generally proceeds as predicted by the enthalpy change $\triangle H$ because

*A) the product $T\triangle S$ may be small compared to $\triangle H$.

B) The product $T\triangle S$ may be large compared to $\triangle H$.

C) The product $T\triangle S$ is equal to $\triangle H$.

D) $\triangle S$ is negative.

33) Which one of the following reactions has a negative $\triangle H$ value?

A) $2Fe_2O_3(s) + 3C(s) \rightarrow 4Fe(s) + 3CO_2(g)$

B) $B(s) \rightarrow B(g)$

C) $CO_2(g) \rightarrow C(s) + O_2(g)$

*D) $2Hg(g) + O_2(g) \rightarrow 2HgO(g)$

13A.2.

34) Warm water has a more rapid rate of evaporation than cold water because

A) it has a higher viscosity
B) its molecules have less attraction for one another
*C) its molecules have a higher average kinetic energy
D) more of its molecules have nearly the same kinetic energy

35) For which one of the following substances is the standard enthalpy of formation, $\triangle H_f$, equal to zero?

A) water ($H_2O(\ell)$)
*B) lead (Pb(s))
C) carbon dioxide ($CO_2(s)$)
D) tin (Sn(g))

13A.3.

1) Consider the following reactions:

$$\text{I} \quad Ag^+(aq) + I^-(aq) \rightarrow AgI(s)$$

$$\text{II} \quad 4Fe(s) + 3O_2(g) \xrightarrow{\text{water}} 2Fe_2O_3(aq)$$

Which one of the following statements best describes the relative rates of the two reactions?

A) II is faster than I
B) I and II are both slow
C) I and II are both fast
*D) I is faster than II

2) For the reaction:

$$Cl_2(\ell) + 2Na(s) \rightarrow 2NaCl(s)$$

The rate of the forward reaction could be increased most by

A) adding more sodium, Na
B) cooling the system
C) removing sodium chloride, NaCl
*D) using gaseous chlorine, $Cl_2(g)$

3) Increasing the concentrations of the reactants in a chemical reaction increases the rate of reaction because the

A) collisions become more effective
*B) frequency of collisions increases
C) average kinetic energy increases
D) activation energy decreases

13A.3.

4) The overall rate of any chemical reaction is most closely related to

A) the overall reaction
B) the number of steps in the reaction
C) the fastest step in the reaction
*D) the slowest step in the reaction

5) In chemical kinetics, the sequence of steps by which chemical reactions may occur is called the reaction

A) velocity
B) rate constant
*C) mechanism
D) order

6) The rate-determining step of a reaction is:

$$A + 2B \rightarrow C$$

the expression for the rate of the reaction, r, is

A) $r = k[A][B]$
*B) $r = k[A][B]^2$
C) $r = k[B]^2$
D) $r = k[A](2[B])^2$

7) Which one of the following statements concerning chemical reactions is true?

A) Increasing the concentration of reacting particles decreases the chance for collisions to occur.
B) A reaction occurs each time reactant molecules collide.
C) A catalyst raises the activation energy barrier.
*D) The slowest step in a reaction mechanism determines the rate of reaction.

8) The equation for the combustion of CH_4 (methane) is:

$$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

If methane is consumed at the rate of 4.48 L/min at STP, the rate of reaction can be stated as

A) 8.96 L/min of O_2 produced

*B) 8.96 L/min of O_2 consumed

C) 8.96 L/min of CO_2 produced

D) 8.96 L/min of CO_2 consumed

9) A certain chemical reaction can be represented by the following equation:

$$2A(g) + B(g) \rightarrow C(g)$$

At a fixed temperature, the initial rate of the reaction was measured for various initial concentrations of A and B. The results were as follows:

Initial concentration (mol/L)		Initial reaction rate (mol/h)
A	B	
0.50	0.50	2.0×10^{-3}
1.0	0.50	8.0×10^{-3}
1.0	1.0	8.0×10^{-3}
1.5	1.0	18×10^{-3}
1.0	1.5	8.0×10^{-3}

The rate law expression for this reaction is

A) Rate = $k[A]\ [B]$

B) Rate = $k[B]^2$

C) Rate = $k[A]^2[B]^2$

*D) Rate = $k[A]^2$

13A.3.

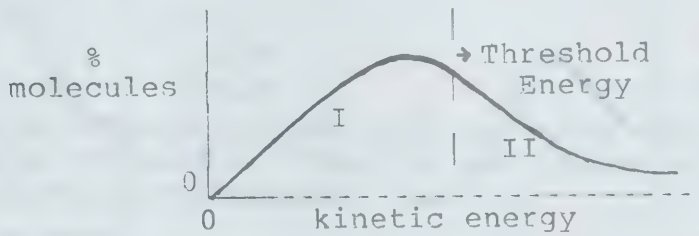
10) The reaction rate expression for a particular exothermic reaction is:

$$\text{Rate} = k[A]^3[B]^2$$

The rate of the reaction can be increased most significantly by

A) decreasing the temperature of the system by 10 K
B) increasing the temperature of the system by 10 K
C) doubling the concentration of B only
*D) tripling the concentration of A only

11) Consider the following graph of the kinetic energy distribution among molecules at temperaure T.



If the temperature were raised to (T + 5) degrees, how would the resulting graph differ from the one above?

A) Both areas I and II would increase.
B) Both areas I and II would decrease.
C) Area I would increase and area II would decrease.
*D) Area I would decrease and area II would increase.

12) The activation energy of a reaction may be defined **qualitatively** as

A) the energy supplied to a catalyst.
B) the energy supplied by the attractive forces between molecules.
C) the heat content of the products minus that of the reactants.
*D) the energy required by reactant molecules to undergo a reaction.

13A.3.

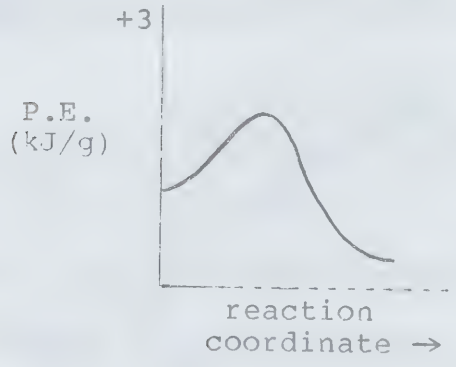
13) Activation energy may be considered as the energy required to

A) make the reactant molecules collide
*B) transform the reactants into an activated complex
C) change the orientation of the reactant molecules
D) increase the vibrational energy of the reactant molecules

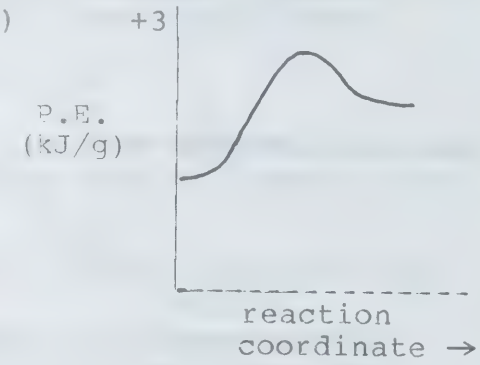
14) Water at 373 K and 101 kPa absorbs 2.3 kJ/g in the process of evaporation. Which of the following potential energy diagrams best represents the reaction?

$$H_2O(\ell) \rightarrow H_2O(g)$$

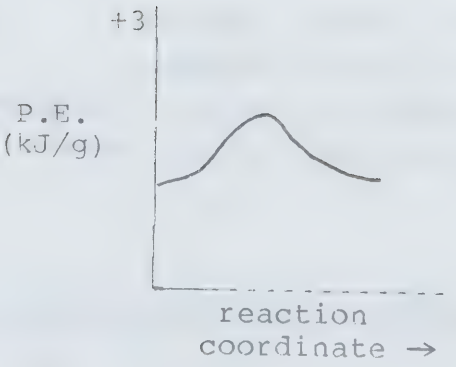
A)



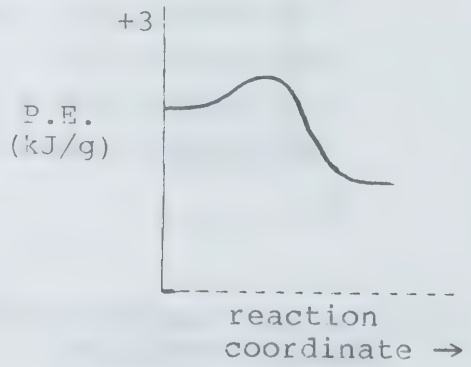
*C)



B)

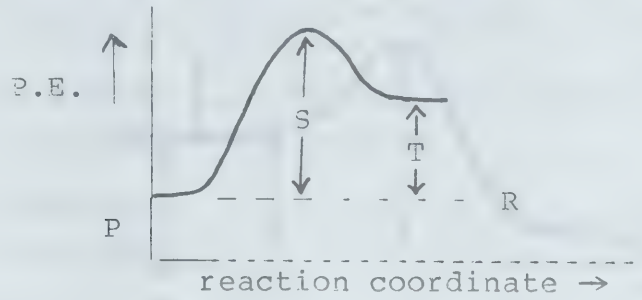


D)



13A.3.

15) In the following potential energy diagram which letter best represents the potential energy of the reactants (prior to reaction)?



*A) P
B) R
C) S
D) T

16) A rise in temperature of 10 K can cause the rate of some reactions to double. This is best explained by which one of the following statements?

A) The average kinetic energy of all the molecules is doubled.
*B) The number of reacting particles with the minimum activation energy is doubled.
C) The activation energy of the reaction is raised.
D) The $\triangle H$, heat of reaction, is lowered.

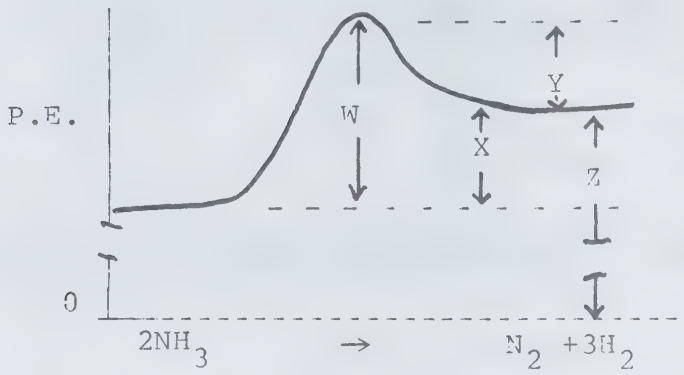
17) The activation energy for an exothermic reaction is 50 kJ/mol. The activation energy for the reverse reaction is

*A) more than 50 kJ/mol
B) less than 50 kJ/mol
C) 50 kJ/mol
D) 0 kJ/mol

13A.3.

18) Given below is a potential energy diagram for the reaction

$$\text{Heat} + 2NH_3(g) = N_2(g) + 3H_2(g)$$



The net heat of reaction corresponds to the interval labelled

A) W
*B) X
C) Y
D) Z

19) The $\triangle H$ for a reaction is 77 kJ/mol and the $\triangle S$ for the reaction is 22 kJ/(mol·K). The temperature at which the reaction becomes spontaneous is

*A) 3.5 K
B) 3.5×10^3 K
C) 55 K
D) 1.7×10^3 K

13A.3.

20) C_4H_{10}(butane) burns in O_2 (oxygen) according to the following balanced equation:

$$C_4H_{10}(g) + 6\tfrac{1}{2}O_2(g) \rightarrow 4CO_2(g) + 5H_2O(g)$$

If butane is consumed at the rate of 2.0 mol/min, the oxygen is consumed at the rate of

A) 6.5 mol/min
B) 9.0 mol/min
*C) 13 mol/min
D) 26 mol/min

21) Which four of the following factors affect the rates of homogeneous chemical reactions?

I	catalyst	V	concentration of reactants
II	entropy	VI	enthalpy
III	free energy	VII	physical nature of reactants
IV	temperature	VIII	equilibrium constant

*A) I, IV, V, VII
B) II, IV, VI, VIII
C) I, IV, VI, VII
D) III, V, VII, VIII

22) Consider the following reactions at 298 K:

$$\text{I} \quad Ag^+(aq) + I^-(aq) \rightarrow AgI(s)$$

$$\text{II} \quad 4Fe(s) + 3O_2(g) \xrightarrow{\text{water}} 2Fe_2O_3(s)$$

Which of the following statements best describes the relative rates of the two reactions?

A) I and II are both slow.
B) I and II are both fast
*C) I is faster than II.
D) II is faster that I.

13A.3.

23) Consider the reaction between $BaCO_3$ (barium carbonate) and an acid:

$$BaCO_3(s) + 2H^+(aq) \rightarrow Ba^{2+}(aq) + H_2O(\ell) + CO_2(g) + \text{heat}$$

Which one of the following changes increases the rate of evolution of CO_2 (carbon dioxide) in this reaction?

A) Decreasing the temperature of the system.

B) Increasing the concentration of Ba^{2+} (barium ions).

C) Adding water to the system.

*D) Finely powdering the $BaCO_3$.

24) At room temperature, and considering only the nature of the reactants, which of the following reactions would be expected to have the greatest rate of reaction?

*A) $Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$

B) $2MnO_4^-(aq) + 16H^+(aq) + 5C_2O_4^{2-}(aq)$

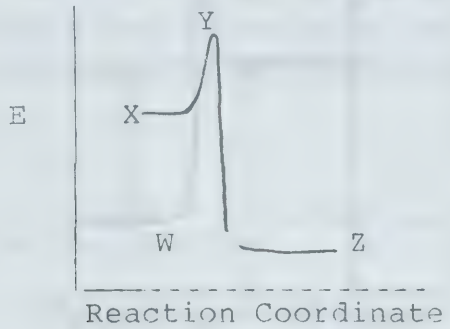
$$\rightarrow 10CO_2(g) + 8H_2O(\ell) + 2Mn^{2+}(aq)$$

C) $C(s) + O_2(g) \rightarrow CO_2(g)$

D) $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$

13A.3.

25) In the following diagram which letter represents the activated complex?



A) X
*B) Y
C) W
D) Z

26) Consider the following reaction mechanism:

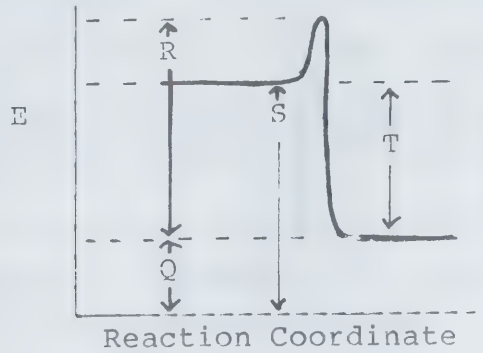
$A + B \rightarrow C$	slow
$C + 2A \rightarrow D + E$	fast
$E \rightarrow F$	fast

The equation for the overall reaction is

A) $A + B \rightarrow F$
*B) $3A + B \rightarrow D + F$
C) $2A + C \rightarrow D + F$
D) $3A + B + C + E \rightarrow C + D + E + F$

13A.3.

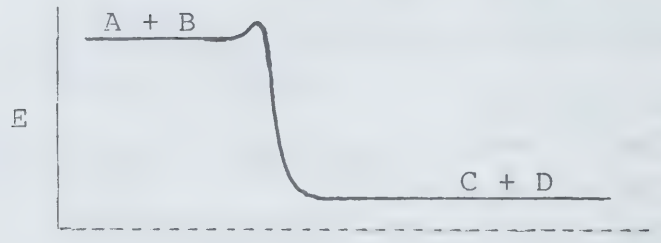
27) In the following potential energy diagram which letter represents $\triangle H$ for the reaction?



A) R
B) Q
C) S
*D) T

28) Given the following potential energy diagram for the reaction:

A + B = C + D



Which of the following is the best interpretation of the graph? The reaction is

*A) exothermic and may be spontaneous
B) exothermic and is not spontaneous
C) endothermic and may be spontaneous
D) endothermic and is not spontaneous

29) A catalyst speeds up a reaction by

A) lowering the $\triangle H$ for the reaction
*B) lowering the activation energy for the reaction
C) raising the kinetic energy of the particles
D) raising the enthalpy of the reactants

30) The mechanism for the reaction

$$H_2 + Br_2 \rightarrow 2HBr$$

suggests that the reaction is a chain reaction. Which of the following steps would terminate the chain?

A) $Br + HBr \rightarrow Br_2 + H\cdot$
B) $H\cdot + Br_2 \rightarrow HBr + Br\cdot$
*C) $2H\cdot \rightarrow H_2$
D) $Br\cdot + H_2 \rightarrow HBr + H\cdot$

34) The rate of the reaction represented by

$$2X(g) + Y(g) \rightarrow Z(g)$$

A) is $k\frac{[Z(g)]}{[X(g)]^2[Y(g)]}$

B) is $k[Z(g)][X(g)][Y(g)]$

C) is $k\frac{[X(g)]^2[Y(g)]}{[Z(g)]}$

*D) cannot be predicted from the equation

35) A $10^{\circ}C$ increase in temperature will double the rate of reaction for:

$$25\ kJ + A \rightarrow B + C$$

because

A) the molecules of A have twice as much kinetic energy

*B) there are twice as many A molecules which are over the threshold energy

C) the system has twice the randomness (entropy)

D) the reaction is endothermic and now has twice as much heat energy

13A.3.

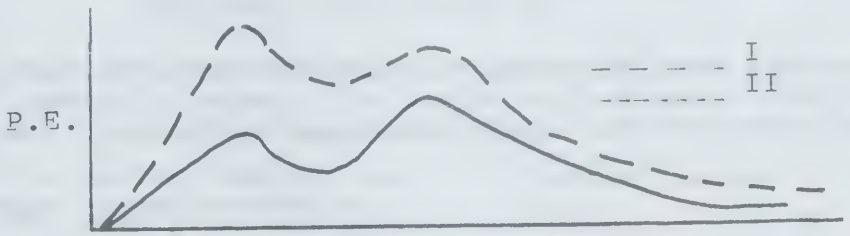
36) An endothermic reaction was found to have the following reaction rate expression:

$$\text{Rate} = k[A]^2[B]^3$$

The **greatest increase** in the rate would be caused by

A) cooling the system
B) removing some of B from the system
C) doubling the concentration of B only
*D) tripling the concentration of A only

37) Substance A reacts to form substance D by the two reaction paths illustrated in the following diagram.



Which one of the following statements regarding Reaction Path I for reaction A → D is true?

A) Path I has a higher overall heat of reaction, $\triangle H$, than Path II.
*B) The rate of reaction by Path II is unaffected by the availability of Path I.
C) Path I represents a faster reaction path than Path II.
D) Path I does not involve the formation of activated complexes.

13A.3.

38) C_4H_{10} (butane) burns in O_2 (oxygen) according to the following balanced equation

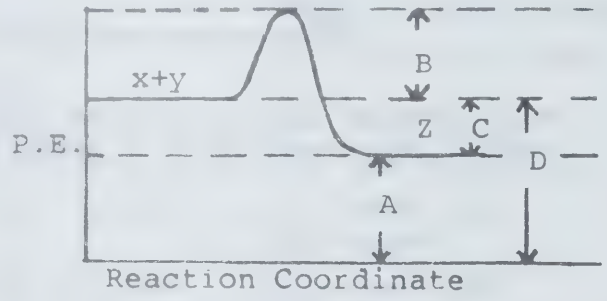
$$C_4H_{10}(g) + \frac{13}{2}O_2(g) \rightarrow 4CO_2(g) + 5H_2O(g)$$

If the butane (at STP) is consumed at the rate of 1.50 L/s, the oxygen (at STP) is consumed at the rate of

A) 1.50 L/s
*B) 9.75 L/s
C) 22.4 L/s
D) 146 L/s

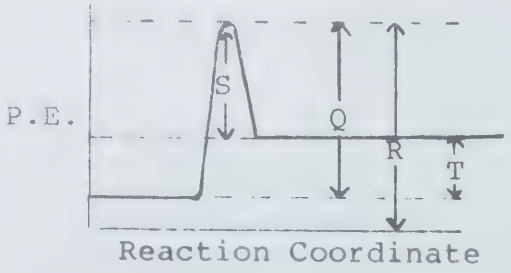
39) Consider the following diagram for the reaction X + Y → Z The addition of a catalyst will cause a change in the magnitude of the energy represented by the interval

A) A
*B) B
C) C
D) D



40) In the following potential energy diagram which letter represents the activation energy for the forward reaction?

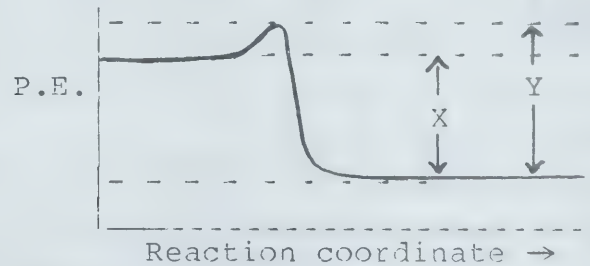
A) R
*B) Q
C) S
D) T



41) Hydrogen ($H_2(g)$) and fluorine ($F_2(g)$) react to form hydrogen fluoride (HF(g)) according to the following equation:

$$\frac{1}{2}H_2(g) + \frac{1}{2}F_2(g) = HF(g) + 270\ kJ$$

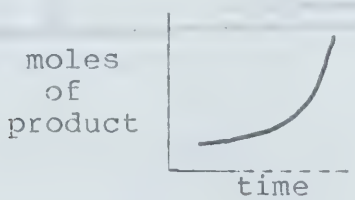
The potential energy diagram for the reaction is given below:

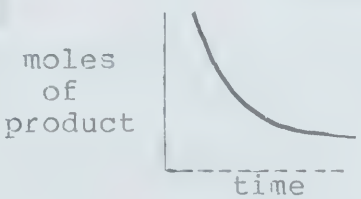


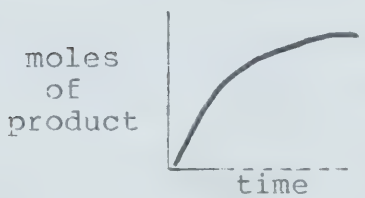
The addition of a catalyst to the reaction vessel would cause

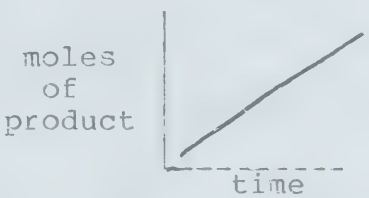
A) both X and Y to become smaller
B) X to become smaller and Y to become larger
*C) X to stay the same and Y to change
D) both X and Y to stay the same

42) In a study of a chemical reaction starting with reactants only, according to the collision theory which of the following graphs best illustrates how the total amount of product varies with time?

A)


C)


*B)


D)


43) Consider the kinetic energy (K.E.) distribution in the following diagram. If the temperature of the system were increased from 20°C to 30°C the % of molecules with KE $\geq E_x$ would

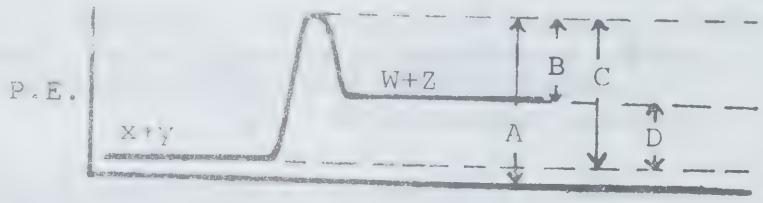
A) decrease by a factor of 10
B) increase by a factor of 10
C) decrease by a factor of 2
*D) increase by a factor of 2

44) On the following potential energy (P.E.) diagram for the reaction:

$$X + Y = W + Z$$

which letter best indicates the potential energy of the activated complex?

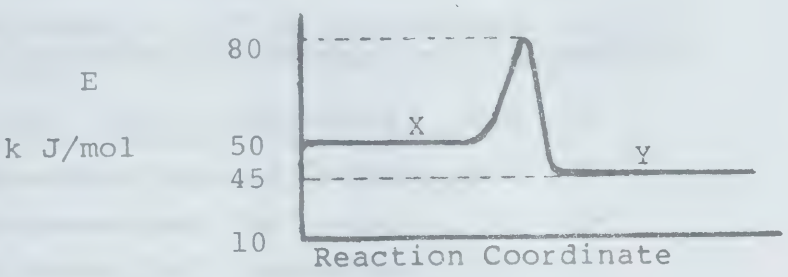
*A) A
B) B
C) C
D) D



13A.3.

45) Substance X reacts to form substance Y by the reaction path in the following diagram.
The overall heat of reaction, $\triangle H$, for the reaction X → Y is

A) -35 kJ/mol
*B) -5 kJ/mol
C) +5 kJ/mol
D) +30 kJ/mol



46) A decrease in entropy occurs when

A) a solid sublimes to a gas
*B) hydrogen and oxygen combine to form water molecules
C) sugar dissolves in water
D) coal burns to carbon dioxide gas and water vapour

13A.4.

1) The following closed system is in dynamic equilibrium:

$$PCl_5(g) = PCl_3(g) + Cl_2(g) \quad \triangle H = 165.9 \text{ kJ/mol}$$

If the temperature of this system is increased at constant pressure, the changes in concentrations would be

A) $[PCl_5]$ increased, $[PCl_3]$ and $[Cl_2]$ decreased
B) $[PCl_5]$ and $[Cl_2]$ increased, $[PCl_3]$ decreased
*C) $[PCl_3]$ and $[Cl_2]$ increased, $[PCl_5]$ decreased
D) $[Cl_2]$ increased, $[PCl_3]$ and $[PCl_5]$ decreased

2) Some ethanol is placed in a closed container of air. The container is heated until the ethanol boils at which time more air is pumped into the flask without altering the temperature. The equilibrium partial pressure of

A) ethanol is increased by the addition of air
B) ethanol is decreased by the addition of air
*C) ethanol is not affected by the addition of air
D) air remains constant

3) NO, O_2 and NO_2 exist in equilibrium according to the following equation:

$$2NO(g) + O_2(g) = 2NO_2(g) \quad \triangle H = -27 \text{ kJ/mol of } NO_2(g)$$

Increasing the temperature of the system at constant pressure will cause

A) an increase in the concentration of $NO_2(g)$
B) a decrease in the volume of the system
*C) an increase in the concentration of $O_2(g)$
D) a decrease in the concentration of NO(g)

4) One of the reactions involved in the production of nitric acid is:

$$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g) + \text{heat}$$

For maximum yield of NO(g)at equilibrium, this reaction should be carried out

A) at high temperature and low pressure
B) at high temperature and high pressure
*C) at low temperature and low pressure
D) at low temperature and high pressure

5) Given the reaction:

$$N_2(g) + 3H_2(g) = 2NH_3(g) + 100\ kJ$$

Which of the following changes will increase the equilibrium concentration of NH_3 (ammonia)?

A) A decrease in the total pressure (at constant temperature)
B) A decrease in the concentration of $H_2(g)$ (at constant volume)
*C) A decrease in the temperature of the reaction chamber (at constant pressure)
D) An increase in the volume of the reaction chamber (at constant temperature)

6) If the equilibrium constant for a reaction is increased by a temperature change, the concentrations of the

A) reactants and products both increase
B) reactants increase and products decrease
*C) products increase and reactants decrease
D) products and reactants both decrease

7) The equilibrium constant expression for the reaction:

$$aA + bB = cC + dD$$

is

A) $K = \frac{a[A]\ b[B]}{c[C]\ d[D]}$

B) $K = [C]^c\ [D]^d$

C) $K = [A]^a [B]^b$

*D) $K = \frac{[C]^c\ [D]^d}{[A]^a\ [B]^b}$

8) Consider the following system:

$$H_2(g) + I_2(g) = 2HI(g)$$

A quantity of pure HI (hydrogen iodide) was placed in an empty vessel at 731 K and the system was allowed to come to equilibrium. An increase in the pressure in the reaction vessel would

A) cause the reaction to proceed to completion

B) cause the production of more H_2

C) cause the production of more HI

*D) cause no effect on the equilibrium

9) Given the following reaction at equilibrium:

$A(g) + 3B(g) = 2C(g) + 2D(g) \quad \triangle H = 21$ kJ/mol

the amount of C in the system can be increased by

A) decreasing the concentration of B at constant temperature and pressure

*B) increasing the temperature at constant pressure

C) adding a catalyst at constant temperature and pressure

D) increasing the pressure at constant temperature

13A.4.

10) The equilibrium constant for a chemical reaction can be changed by

*A) increasing the temperature
B) adding a catalyst
C) increasing the concentration of the products
D) increasing the concentration of the reactants

11) In which of the following reactions does the tendency towards maximum **entropy** favour production of the products?

A) $CO_2(g) = CO_2(s)$

*B) $H_2O(\ell) = H_2O(g)$

C) $Ag^+(aq) + NO_3^-(aq) = AgNO_3(s)$

D) $4Fe(s) + 3O_2(g) = 2Fe_2O_3(s) + 1680\ kJ$

12) For the reaction:

$$2A(g) = 3B(g) + C(g)$$

the equilibrium constant expression is

A) $\frac{[A]^2}{[C][B]^3}$

B) $\frac{(2[A])^3}{[C](3[B])^2}$

*C) $\frac{[C][B]^3}{[A]^2}$

D) $\frac{[C](3[B])^3}{(2[A])^2}$

13A.4.

13) For the following reaction the equilibrium concentrations at a given temperature are $[NO(g)] = 1.0$ mol/L, $[O_2(g)] = 6.0$ mol/L, $[NO_2(g)] = 2.0$ mol/L:

$$2NO(g) + O_2(g) = 2NO_2(g)$$

The value of K_{eq} for the system at this temperature is

A) 0.33
*B) 0.67
C) 1.5
D) 9.0

14) 0.80 mol/L solutions of each of A and B are mixed. They react slowly, producing C and D according to the following equation:

$$A(aq) + B(aq) = C(aq) + 2D(aq)$$

If the equilibrium concentration of C is 0.60 mol/L, the equilibrium concentration of D is

*A) 0.20 mol/L
B) 0.40 mol/L
C) 0.60 mol/L
D) 0.80 mol/L

15) $H_2(g)$, $Cl_2(g)$ and HCl(g) exist in equilibrium, at a fixed temperature, according to the equation:

$$2HCl(g) = H_2(g) + Cl_2(g) \qquad K_{eq} = 0.020$$

In order to have 0.020 mol of $Cl_2(g)$ (chlorine gas) at equilibrium, how many moles of HCl(g) (hydrogen chloride gas) should be initially placed in an empty 1.0 L container?

A) 0.04
B) 0.14
*C) 0.18
D) 0.22

16) For the following equilibrium what is the effect of adding $O_2(g)$ (oxygen) to the system, at constant volume and temperature?

$$2NO(g) + O_2(g) = 2NO_2(g)$$

A) No change in the concentrations of NO(g) or $NO_2(g)$
B) A decrease in the concentration of $NO_2(g)$
C) An increase in the concentration of NO(g)
*D) An increase in the concentration of $NO_2(g)$

17) H_2 (hydrogen) can be produced from CH_4 (methane) by the following reaction

$$CH_4(g) + 2H_2O(g) = CO_2(g) + 4H_2(g) \qquad \triangle H = -163.8 \text{ kJ}$$

Assuming equilibrium can be rapidly attained under all conditions, the greatest degree of conversion of methane to hydrogen would be expected at

A) low temperature and high pressure
*B) low temperature and low pressure
C) high temperature and low pressure
D) high temperature and high pressure

13A.4.

18) Consider the following system:

$$H_2(g) + I_2(g) = 2HI(g)$$

A quantity of pure HI (hydrogen iodide) was placed in an empty flask at 731 K and allowed to come to equilibrium. At equilibrium the flask contained 1.00 mol of $H_2(g)$. If the equilibrium constant for the reaction, as written, is 50.6, how many moles of HI(g) are present in the flask at equilibrium?

*A) 7.1
B) 2.00
C) 48.1
D) 48.6

19) Which of the following operations will have the greatest effect on an equilibrium constant?

A) Adding a catalyst at constant temperature.
B) Increasing the pressure at constant temperature.
*C) Increasing the temperature at constant pressure.
D) Increasing the concentration of products at constant temperature.

20) The equilibrium law expression for the following balanced equation is:

$$2Q(g) + P(g) = 3X(g) + Y(g)$$

A) $\dfrac{[3X(g)][Y(g)]}{[2Q(g)][P(g)]}$

B) $\dfrac{[Q(g)]^2[P(g)]}{[X(g)]^3[Y(g)]}$

C) $\dfrac{[X(g)]^3 + [Y(g)]}{[Q(g)]^2 + [P(g)]}$

*D) $\dfrac{[X(g)]^3[Y(g)]}{[Q(g)]^2[P(g)]}$

13A.4.

21) In the system:

$$A(aq) + 2B(aq) = C(aq)$$

the concentrations of A, B, and C in the equilibrium mixture were found to be 2 mol/L, 4 mol/L and 1 mol/L respectively. The equilibrium constant for this reaction is

*A) $\frac{1}{32}$

B) $\frac{1}{16}$

C) 8

D) 16

22) Consider a stoppered flask containing benzene, iodine and air which has reached equilibrium at 298 K. Which one of the following statements concerning the system at equilibrium is true?

A) The pressure of the vapour phase is not constant.
*B) Evaporation of the liquid continues.
C) The colour of the liquid phase will change.
D) The mass of solid iodine is not constant.

23) Consider the equilibrium represented by the following equation:

$$\text{Heat} + H_2O(\ell) = H_2O(g) \qquad [T = 373\ K,\ P = 101\ kPa]$$

The reverse reaction is favoured when

A) the total volume of the system is increased
B) boiling water is added to the system
C) water vapour is removed from the system
*D) the system is cooled

13A.4.

24) Consider the reaction:

$$2X(g) + Y(g) = 2Z(g) \text{ (at 298 K)}$$

The initial concentrations in the system are:

$[X] = 1.5 \times 10^{-2}$ mol/L, $[Y] = 2.0 \times 10^{-1}$ mol/L, $[Z] = 0$

After the system comes to equilibrium, $[Z] = 2.1 \times 10^{-3}$ mol/L. What are the equilibrium concentrations in mol/L of X and Y at 298 K?

*A) $[X] = 1.3 \times 10^{-2}$, $[Y] = 2.0 \times 10^{-1}$

B) $[X] = 1.7 \times 10^{-2}$, $[Y] = 2.1 \times 10^{-1}$

C) $[X] = 2.9 \times 10^{-2}$, $[Y] = 2.0 \times 10^{-2}$

D) $[X] = 2.9 \times 10^{-1}$, $[Y] = 4.0 \times 10^{-1}$

25) Consider the following reaction:

$$3X + Y = 2W + Z$$

At 298 K the equilibrium concentrations of X, Y, W and Z were found to be 7.0 mol/L, 6.0 mol/L, 5.0 mol/L and 7.0 mol/L, respectively. When the temperature was raised to 313 K the new equilibrium concentration of Z was found to be 6.0 mol/L. What are the new equilibrium concentrations of the other three substances?

A) [X] = 8.0 mol/L, [Y] = 7.0 mol/L, [W] = 4.0 mol/L

B) [X] = 4.0 mol/L, [Y] = 5.0 mol/L, [W] = 7.0 mol/L

*C) [X] = 10.0 mol/L, [Y] = 7.0 mol/L, [W] = 3.0 mol/L

D) [X] = 10.0 mol/L, [Y] = 7.0 mol/L, [W] = 4.0 mol/L

26) An example of a macroscopic property is the

A) angle of collision between molecules

B) shape of the CO_2 molecule

*C) colour of NO_2 gas

D) relative atomic mass of He

13A.4.

27) Which of the following systems, at room temperature and pressure, would be described as being in dynamic equilibrium?

*A) a stoppered flask containing air, water and water vapour

B) an open flask containing air, water and water vapour

C) a stoppered flask containing $H_2(g)$, $O_2(g)$ and $H_2O(g)$

D) an open flask containing sodium metal (Na)

28) Acetic acid ionizes according to the equation

$$CH_3COOH(aq) = H^+(aq) + CH_3COO^-(aq)$$

If sodium acetate (CH_3COONa) is added to the acid solution, which of the following will occur?

*A) the hydrogen ion concentration will decrease

B) the solution will become more acidic

C) a precipitate will form

D) more acetic acid will ionize

29) For the following system in equilibrium, increasing the temperature of the system at constant pressure will cause

$$2NO(g) + O_2(g) = 2NO_2(g) \qquad \triangle H = -27 \text{ kJ}$$

A) An increase in the concentration of $NO_2(g)$.

*B) An increase in the concentration of $O_2(g)$.

C) A decrease in the volume of the system.

D) A decrease in the concentration of $NO(g)$.

13A.4.

30) Consider the following system:

$$H_2(g) + I_2(g) = 2HI(g)$$

A quantity of pure HI(g) (hydrogen iodide) was placed in an empty flask at 731 K, stoppered and allowed to come to equilibrium. If a sample of $H_2(g)$ is added to the system at equilibrium

*A) [HI] increases and $[I_2]$ decreases.

B) [HI] decreases and $[I_2]$ increases.

C) [HI] increases and $[I_2]$ does not change.

D) [HI] increases and $[I_2]$ increases.

31) In the reaction A(aq) + 2B(aq) = C(aq), the equilibrium concentrations of A, B, and C were found to be 2, 4, and 1 mol/L respectively. The equilibrium constant for the reaction is

A) 16

B) 32

C) $\frac{1}{16}$

*D) $\frac{1}{32}$

13A.4.

32) Consider the equations:

I) $Ni^{2+}(aq) + Co^{2+}(aq) = NiCo^{4+}(aq)$

II) $Ni^{2+}(aq) + Co^{2+}(aq) = NiCo^{4+}(s)$

III) $Ni(s) + Co^{2+}(aq) = Ni^{2+}(aq) + Co(s)$

IV) $Ni^{2+}(aq) + Co(aq) = Ni(s) + Co^{2+}(aq)$

$$K_x = \frac{[Ni^{2+}(aq)]}{[Co^{2+}(aq)]}$$

K_x is the equilibrium constant for equation

A) I
B) II
*C) III
D) IV

33) Reactants A and B are mixed. They react slowly, producing C and D according to the following equation:

$$A(g) + B(g) = C(g) + 2D(g)$$

The equilibrium constant expression for the reaction is

A) $K = \frac{[A][B]}{[C][D]^2}$

B) $K = \frac{[A][B]}{[C][D]}$

C) $K = \frac{[C][D]}{[A][B]}$

*D) $K = \frac{[C][D]^2}{[A][B]}$

13A.4.

34) Consider the reaction:

$$2SO_2(g) + O_2(g) = 2SO_3(g)$$

The equilibrium concentrations at a given temperature are as follows:

$[SO_2] = 0.20$ mol/L, $[O_2] = 0.10$ mol/L, $[SO_3] = 0.50$ mol/L

What is the value of the equilibrium constant for this system?

A) 0.55
B) 5.5
C) 25
*D) 63

35) Consider the following systems:

I A container is half-filled with alcohol, stoppered and allowed to stand for several days.

II Crystals of $KMnO_4$ (potassium permanganate) are dissolved in water until the water becomes a uniform purple colour and a constant amount of crystal remains in the container.

III An ice cube floats in a beaker of water in a normal classroom.

IV Moth balls are placed in an empty flask which is then stoppered and allowed to stand for several days.

Which of the above are in an equilibrium state?

A) I and II only
*B) I, II and IV only
C) II and III only
D) I, II, III, and IV

13A.4.

36) 1.5 g samples of magnesium ribbon are each added to 1 L of acetic acid (2 mol/L) and 1 L of hydrochloric acid (2 mol/L). Which of the following best describes the reactions that occur in the acid solutions?

A) Both the rate of reaction and the mass of Mg consumed are the same.

B) Both the rate of reaction and the mass of Mg consumed are different.

*C) The rate of reaction is different but the mass of Mg consumed is the same.

D) The rate of reaction is the same but the mass of Mg consumed is different.

37) What is the equilibrium constant, at 298 K, for the following reaction, given that the equilibrium concentrations of X, Y, W and Z at 298 K are 7.0 mol/L, 6.0 mol/L, 5.0 mol/L and 7.0 mol/L, respectively?

$$3X + Y = 2W + Z$$

*A) 0.085

B) 0.56

C) 1.8

D) 12

38) P and Q react to form X and Y according to the equation

$$2P(aq) + Q(aq) = 2X(aq) + Y(aq)$$

In a closed system the initial concentrations of P and Q were each 1.2 mol/L; the initial concentrations of X and Y were 0 mol/L. When equilibrium was established the concentration of X was 0.80 mol/L. The equilibrium concentrations in mol/L are

*A) P: 0.40, Q: 0.80, Y: 0.40

B) P: 0.40, Q: 0.80, Y: 1.6

C) P: 0.40, Q: 0.40, Y: 0.40

D) P: 0.80, Q: 0.40, Y: 1.6

13A.4.

39) The value of the equilibrium constant is 50 for the reaction:

$$H_2(g) + I_2(g) = 2HI(g)$$

If $[H_2]$ = 1.2 mol/L and $[I_2]$ = 0.60 mol/L at equilibrium, the equilibrium concentration of HI(g) is

A) 1.4×10^{-2} mol/L
B) 2.4 mol/L
*C) 6.0 mol/L
D) 36 mol/L

13A.5.

1) The solubility in water of Ag_2S (silver sulfide) is 1.3×10^{-6} mol/L at 293 K. The K_{sp} value of Ag_2S is

A) 1.6×10^{-18}
*B) 8.8×10^{-18}
C) 1.7×10^{-12}
D) 1.1×10^{-3}

2) The solubility of a gas in a liquid will be greatest at

A) low temperature and low pressure
B) high temperature and low pressure
C) high temperature and high pressure
*D) low temperature and high pressure

3) At 293 K the solubility of KCl (potassium chloride) is 34.0 g per 100 g of water. What mass of KCl must be dissolved in water to form 200 g of saturated solution at 293 K?

A) 39.3 g
*B) 50.7 g
C) 68.0 g
D) 103 g

4) A strong electrolyte is a substance that

A) reacts with water to produce a strong acid
B) is very soluble in water
*C) forms aqueous solutions which are good conductors of electricity
D) is metallic and has loosely bound electrons

13A.5.

5) X is an element in Group I and E is an element in Group VII of the Periodic Table. The compound formed when X and E react will most likely

A) be a solid with high electrical conductivity
B) have a melting point below 373 K
*C) dissolve in water to form a strong electrolyte
D) be readily soluble in a non-polar solvent

6) Which one of the following aqueous solutions, at $25^{\circ}C$, is the poorest electrical conductor?

A) 1.0 mol/L HCl (hydrochloric acid) solution
*B) 1.0 mol/L C_2H_5OH (ethyl alcohol) solution
C) 1.0 mol/L NaCl (sodium chloride) solution
D) 1.0 mol/L KOH (potassium hydroxide) solution

7) Which of the following compounds has the greatest solubility in water at 298 K?

*A) ammonium nitrate
B) copper(II) sulfide
C) silver carbonate
D) lead(II) chloride

8) A saturated solution of CH_3COOAg (silver acetate) contains 2×10^{-3} mol of Ag^+ per litre of solution.
The K_{sp} (solubility product) of silver acetate is

A) 2×10^{-6}
*B) 4×10^{-6}
C) 1×10^{-3}
D) 2×10^{-3}

13A.5.

9) The K_{sp} (solubility product) for AgCl (silver chloride) in water at 298 K is 1.7×10^{-10}.
What is the solubility of AgCl in mol/L?

A) 1.7×10^{-10}
B) 2.4×10^{-10}
C) 3.4×10^{-7}
*D) 1.3×10^{-5}

10) Given that K_{sp} for $AgIO_3$ (silver iodate) is 4.0×10^{-8}, the concentration of $Ag^+(aq)$ (silver ion) in a saturated solution is

A) 1.6×10^{-15} mol/L
B) 4.0×10^{-8} mol/L
*C) 2.0×10^{-4} mol/L
D) 4.0×10^{-4} mol/L

11) In acidic solutions the pH is

A) 10^{-7}
*B) <7
C) 7
D) >7

12) Given the following solubility product information:

Silver bromide (AgBr)	$K_{sp} = 7.7 \times 10^{-13}$
Silver chloride (AgCl)	$K_{sp} = 1.6 \times 10^{-10}$
Silver hydroxide (AgOH)	$K_{sp} = 2.0 \times 10^{-8}$
Silver iodide (AgI)	$K_{sp} = 1.5 \times 10^{-16}$

The order of increasing solubility of the four compounds, from left to right, is

A) AgI, AgBr, AgOH, AgCl
B) AgOH, AgCl, AgBr, AgI
*C) AgI, AgBr, AgCl, AgOH
D) AgBr, AgCl, AgI, AgOH

13) The value of the solubility product of AX is 1.6×10^{-5} at 353 K. AX dissolves in water according to the following equation:

$$AX(s) = A^{+}(aq) + X^{-}(aq)$$

The solubility of AX is

A) 0.80×10^{-5} mol/L
B) 1.6×10^{-5} mol/L
*C) 4.0×10^{-3} mol/L
D) $3.2 \times 10^{-2.5}$ mol/L

14) The solubility product, K_{sp}, for AgCl (silver chloride) is 1.56×10^{-10}. The solubility of AgCl in water is

A) 7.80×10^{-11} mol/L
B) 1.56×10^{-10} mol/L
C) 2.12×10^{-10} mol/L
*D) 1.25×10^{-5} mol/L

13A.5.

15) Consider the system

$$MX\ (s) = M^{2+}(aq) + X^{2-}(aq).$$

If the K_{sp}, solubility product, for MX is 9.8×10^{-7}, then the solubility of MX in water is

A) 4.90×10^{-7} mol/L
B) 9.80×10^{-7} mol/L
C) 1.96×10^{-6} mol/L
*D) 9.90×10^{-4} mol/L

16) What is the concentration of silver ions in saturated aqueous silver bromide solution if the solubility product constant of silver bromide is 3.3×10^{-13}?

A) 1.7×10^{-13} mol/L
B) 6.6×10^{-7} mol/L
*C) 5.7×10^{-7} mol/L
D) 4.3×10^{-5} mol/L

17) Which one of the following pairs of substances will form a precipitate in dilute aqueous solution?

A) KNO_3 and NaCl
B) $MgCl_2$ and $Zn(NO_3)_2$
C) NH_4Cl and $Fe_2(SO_4)_3$
*D) $Cu(NO_3)_2$ and Na_2CO_3

13A.5.

18) One litre of a solution contains 10^{-5} mol of each of I^-, Br^-, and OH^- ions. If 10^{-3} mol of Ag^+ is added to the solution.

$$AgI: K_{sp} = 1.5 \times 10^{-16}$$
$$AgBr: K_{sp} = 7.7 \times 10^{-13}$$
$$AgOH: K_{sp} = 1.6 \times 10^{-5}$$

The resultant precipitate would contain

A) AgOH, AgI and AgBr
*B) AgI and AgBr only
C) AgOH and AgBr only
D) AgOH only

19) The solubility of $SrSO_4$ (strontium sulfate) at room temperature is 8.0×10^{-4} mol/L. The K_{sp} of $SrSO_4$ is

A) 1.6×10^{-9}
B) 1.6×10^{-7}
*C) 6.4×10^{-7}
D) 4.0×10^{-4}

20) The solubility of strontium sulfate ($SrSO_4$) at room temperature is 8.0×10^{-4} mol/L. The solubility product of strontium sulfate is

A) 1.6×10^{-8}
*B) 6.4×10^{-7}
C) 8.0×10^{-4}
D) 4.0×10^{-4}

13A.5.

21) Consider the equation:

$$AgIO_3(s) = Ag^+(aq) + IO_3^-(aq).$$

The K_{sp} of $AgIO_3$ (silver iodate) is 2.10×10^{-8}. The minimum number of moles of $AgNO_3$ (silver nitrate) that must be added to 1.00 L of 2.00×10^{-3} mol/L KIO_3 (potassium iodate) solution in order to cause a precipitate to form is

A) 1.05×10^{-8}
*B) 1.05×10^{-5}
C) 2.10×10^{-5}
D) 3.00×10^{-3}

22) A saturated solution of CH_3COOAg (silver acetate) contains 2×10^{-3} mol of Ag^+ per litre of solution. The K_{sp} for silver acetate is

A) 2×10^{-6}
*B) 4×10^{-6}
C) 2×10^{-3}
D) 1×10^{-3}

23) $BaCO_3$ (barium carbonate) dissolves slightly in water according to the following equation:

$$BaCO_3(s) = Ba^{2+}(aq) + CO_3^{2-}(aq)$$

At 298 K, $K_{sp} = [Ba^{2+}][CO_3^{2-}] = 8.1 \times 10^{-9}$. In a saturated solution at 298 K, the concentration of $Ba^{2+}(aq)$ ion is

A) 9.0×10^{-10} mol/L
B) 8.1×10^{-9} mol/L
C) 16×10^{-9} mol/L
*D) 9.0×10^{-5} mol/L

13A.5.

24) At 303 K, 9.00 g of KCl(s) (potassium chloride) will exactly saturate 25.0 mL of water. What is the solubility of KCl in grams per litre of water?

A) 225
B) 265
C) 278
*D) 360

25) The K_{sp} for $CaF_2(s)$ (calcium fluoride) is 1.7×10^{-10}. The highest concentration of Ca^{2+} ion that can be present in a solution containing 0.010 mol/L F^- ion is

A) 8.5×10^{-9} mol/L
B) 1.7×10^{-8} mol/L
C) 4.3×10^{-7} mol/L
*D) 1.7×10^{-6} mol/L

26) Which one of the following is the correct equation for the dissociation of K_2SO_4 (potassium sulfate) in water?

A) $K_2SO_4(s) \rightarrow K^{2+}(aq) + SO_4^-(aq)$

B) $K_2SO_4(s) \rightarrow 2K^-(aq) + SO_4^{2+}(aq)$

C) $K_2SO_4(s) \rightarrow K^+(aq) + 2SO_4^{2-}(aq)$

*D) $K_2SO_4(s) \rightarrow 2K^+(aq) + SO_4^{2-}(aq)$

27) The number of moles of NaCl (sodium chloride) present in 150 mL of 0.0200 mol/L NaCl(aq) is

A) 1.33×10^{-4}
*B) 3.00×10^{-3}
C) 3.00
D) 0.130

13A.5.

28) The solubility product of AgCl (silver chloride) is 1.4×10^{-4} at 373 K. The solubility of AgCl in water at 373 K is

A) 0.60×10^{-4} mol/L
B) 0.70×10^{-4} mol/L
C) 1.4×10^{-4} mol/L
*D) 1.2×10^{-2} mol/L

29) The correct K_{sp} expression for Ag_2S (silver sulphide) is

A) $Ksp = [Ag^+]^2[S^{2-}]^2$
B) $Ksp = ([Ag^+])[S]$
*C) $Ksp = [Ag^+]^2[S^{2-}]$
D) $Ksp = 2[Ag^+]^2[S^{2-}]$

30) 400 mL of a saturated solution of TlCl (thallium(I) chloride) in water at 298 K was removed from the container and evaporated to dryness. The residue had a mass of 1.20 g. The relative atomic masses of Tl and Cl are 204 and 35.5 respectively. The solubility product constant, Ksp, for TlCl, at 298 K is

A) 2.5×10^{-5}
*B) 1.6×10^{-4}
C) 2.5×10^{-2}
D) 1.3×10^{2}

13A.5.

31) What is the solubility of a salt in g per 100 g of solvent if the salt forms a saturated solution containing 12.0 g of salt in 30.0 g of water at 298 K?

A) 17.1
B) 28.6
C) 36.0
*D) 40.0

32) The solubility of $SrSO_4$ (strontium sulfate) at room temperature is 8.0×10^{-4} mol/L. The calculated Ksp, solubilit product, for $SrSO_4$ is

A) 6.4×10^{-9}
B) 1.6×10^{-7}
*C) 6.4×10^{-7}
D) 4.0×10^{-4}

33) A sample of H_2SO_4 (sulfuric acid) has a density of 1.80 g/mL and contains 80.0% solute by mass. The mass of solute present in 150 mL of this sample is

A) 83.3 g
B) 120 g
*C) 216 g
D) 270 g

34) A solution of NaCl (sodium chloride) has a concentration of 1.8 mol/L and has a density of 1.2 g/mL. The percentage by mass of NaCl in this solution is

A) 4.2%
B) 5.9%
*C) 8.8%
D) 9.6%

13A.5.

35) What mass of I_2 (iodine) is present in 3.0×10^2 g of a solution of iodine in alcohol, if the solution is 3.0% solute by mass?

A) 6.0 g
*B) 9.0 g
C) 12 g
D) 30 g

36) Salts in aqueous solution are always

*A) highly ionized
B) highly associated
C) weakly ionized
D) acidic

37) Which one of the following terms best describes compounds which, in solution, conduct electricity?

A) non-electrolyte
*B) electrolyte
C) hydrate
D) anhydride

38) Sugar dissolved in water

*A) does not conduct electricity
B) forms hydrated ions
C) conducts electricity
D) donates protons to H_2O

13A.5.

39) Which one of the following is a salt?

*A) KNO_3
B) CH_3COOH
C) $Ca(OH)_2$
D) $H_2C_2O_4$

40) Which of the following substances is **LEAST** soluble in water at room temperature?

*A) silver iodide (AgI); Ksp = 8.5×10^{-17}
B) silver bromide (AgBr); Ksp = 5.0×10^{-13}
C) silver iodate ($AgIO_3$); Ksp = 3.1×10^{-8}
D) calcium sulfate ($CaSO_4$); Ksp = 2.4×10^{-5}

41) Equal volumes of 0.10 mol/L solutions of two salts are mixed. In which case would a bright yellow precipitate be formed?

A) silver nitrate ($AgNO_3(aq)$) and sodium chloride ($NaCl(aq)$)
B) calcium nitrate ($Ca(NO_3)_2(aq)$) and potassium sulphate ($K_2SO_4(aq)$)
*C) lead(II) nitrate ($Pb(NO_3)_2(aq)$) and potassium chromate ($K_2CrO_4(aq)$)
D) zinc nitrate ($Zn(NO_3)_2(aq)$) and sodium sulphide ($Na_2S(aq)$)

13A.5.

42) The solubility in water of lead chromate ($PbCrO_4$) at 298 K is 7.0×10^{-5} g/L. How much lead chromate will dissolve in 50 mL of water to form a saturated solution?

A) 2.4×10^{-7} g
*B) 3.5×10^{-6} g
C) 6.8×10^{-6} g
D) 1.4×10^{-3} g

43) A solution of calcium nitrate ($Ca(NO_3)_2$) is 40% solute by mass. The mass of calcium nitrate present in 85 g of solution is

A) 17 g
*B) 34 g
C) 40 g
D) 85 g

44) What is the concentration in mol/L, of a sodium hydroxide (NaOH) solution containing 10.0 g of sodium hydroxide in 200 mL of solution?

A) 0.0500 mol/L
B) 0.250 mol/L
C) 0.800 mol/L
*D) 1.25 mol/L

45) The Ksp, solubility product, for strontium carbonate ($SrCO_3$) is 1.6×10^{-9}. The solubility of $SrCO_3$ in a 0.10 mol/L solution of sodium carbonate (Na_2CO_3) is

A) 1.6×10^{-10} mol/L
*B) 1.6×10^{-8} mol/L
C) 1.2×10^{-5} mol/L
D) 4.0×10^{-4} mol/L

13A.5.

46) Equal volumes of 2×10^{-3} mol/L aqueous solutions of strontium nitrate ($Sr(NO_3)_2$) and sodium sulphate (Na_2SO_4) are mixed together. Which one of the following statements concerning the solution formed by mixing the two original solutions is true?

*A) $[Na^+] = 2 \times 10^{-3}$ mol/L
B) $[Sr^{2+}] = 4 \times 10^{-3}$ mol/L
C) $[Na^+] = 1 \times 10^{-3}$ mol/L
D) $[Sr^{2+}] = 2 \times 10^{-3}$ mol/L

47) Barium chloride solution is added to an unknown solution and a white precipitate is formed. The solution most probably contains aqueous

*A) Ag^+
B) Na^+
C) Zn^{2+}
D) Al^{3+}

48) In aqueous solution, salts exist in the form of

A) molecules
B) atoms
*C) ions
D) crystals

49) The solubility of $NH_4Cl(s)$ (ammonium chloride) is 41.0 g per 100 g of water at 303 K. How much $NH_4Cl(s)$ would be dissolved in 150 g of saturated solution at 303 K?

A) 27.3 g
B) 41.0 g
*C) 43.6 g
D) 61.5 g

50) An ionic solid A_2C_3 dissociates in water according to the following equation:

$$A_2C_3(s) = 2A^{3+}(aq) + 3C^{2-}(aq)$$

If the solubility of A_2C_3 is 2.0 mol/L, the Ksp (solubility product) of A_2C_3 is

A) 3.5×10^{-3}
B) 24
C) 32
*D) 3.5×10^{3}

51) A clear solution is produced by mixing 1.0 L of 0.50 mol/L NaCl (sodium chloride) with 1.5 L of 0.40 mol/L $MgCl_2$ (magnesium chloride) and 0.50 L of water. The resultant concentration of chloride ion is

A) 0.37 mol/L
B) 0.45 mol/L
*C) 0.57 mol/L
D) 0.60 mol/L

52) The substance which forms the best conducting solution in water is

A) CCl_4 (carbon tetrachloride)
*B) NaCl (sodium chloride)
C) SiC (silicon carbide)
D) O_2 (oxygen)

53) What is the colour of the precipitate formed by mixing solutions containing $Ba^{2+}(aq)$ and $SO_4^{2-}(aq)$?

A) blue
*B) white
C) yellow
D) green

54) Given that the solubility product constant of AgBr (silver bromide) is 3.30×10^{-13}, the concentration of silver ions in a saturated AgBr solution is

A) 1.65×10^{-13} mol/L
B) 3.30×10^{-13} mol/L
C) 1.82×10^{-7} mol/L
*D) 5.74×10^{-7} mol/L

55) A solution of NaOH (sodium hydroxide) has a density of 1.1 g/mL and contains 30% solute by mass. The mass of the solute present in 75 mL of this solution is

A) 18 g
B) 23 g
*C) 25 g
D) 75 g

56) At 298 K, 6.00 g of $KNO_3(s)$ (potassium nitrate) will exactly saturate 15.0 g of water. What is the solubility of KNO_3 in g per 100 g of water?

A) 0.900
B) 21.0
C) 28.6
*D) 40.0

57) 45.0 g of solution contains 40.0 g of solvent. The percent by mass of solute in this solution is

A) 6.00%
*B) 11.1%
C) 12.5%
D) 88.9%

58) The Ksp for $MgCO_3$ (magnesium carbonate) is 2.5×10^{-5}. Equal volumes of the following solutions are mixed

I) 0.0002 mol/L Mg^{2+} and 0.0002 mol/L CO_3^{2-}

II) 0.002 mol/L Mg^{2+} and 0.002 mol/L CO_3^{2-}

III) 0.02 mol/L Mg^{2+} and 0.02 mol/L CO_3^{2-}

IV) 0.2 mol/L Mg^{2+} and 0.2 mol/L CO_3^{2-}

A precipitate of $MgCO_3$ forms in

A) all of the mixtures
B) I and II only
C) II, III and IV only
*D) III and IV only

59) A 0.1 mol/L aqueous solution of a weak monoprotic acid has a hydrogen ion concentration of 0.0010 mol/L. The value of the ionization constant, K_a, for this acid is

A) 10^{-6}
B) 10^{-2}
C) 10^{-3}
*D) 10^{-5}

60) An Arrhenius acid

A) is an ionic compound which contains hydrogen atoms
B) produces an aqueous solution whose pH is greater than 7
C) is always a neutral molecule
*D) releases hydrogen ions in aqueous solution

13A.5.

61) Acetic acid dissociates according to the following equation:

$$CH_3COOH(aq) = H^+(aq) + CH_3COO^-(aq)$$

If CH_3COONa (sodium acetate) is added to the acid, which of the following will occur?

A) The concentration of acetic acid will decrease.
*B) The hydrogen ion concentration will decrease.
C) A precipitate will form.
D) More acetic acid will ionize.

13A.6.

1) The number of moles of atoms in 4.0×10^{60} atoms is

A) 1.5×10^{-37}
B) 6.6×10^{-36}
*C) 6.6×10^{36}
D) $2.4 \times \times 10^{84}$

2) How many moles of NaOH (sodium hydroxide) are present in 65.00 g of NaOH(s)?

A) 0.6500
B) 1.250
*C) 1.625
D) 3.325

3) Aqueous solutions of strong acids and of strong bases

*A) are good conductors of electricity
B) turn litmus red
C) have a pH of 7
D) react with copper to produce hydrogen gas

4) A 12 mol/L solution of $C_2H_4O_2$ (acetic acid) is best described as

A) a concentrated strong acid
B) a dilute strong acid
*C) a concentrated weak acid
D) a dilute weak acid

5) The correct expression for the ion product constant of water (K_w) at 298 K is

*A) $K_w = [H^+]\ [OH^-]$

B) $K_w = \dfrac{[H_2O]}{[H^+]\ [OH^-]}$

C) $K_w = [H^+] + [OH^-]$

D) $K_w = \dfrac{[H^+]\ [OH^-]}{[H_2O]}$

6) HI (hydroiodic acid) is a very strong acid. In an aqueous solution of 0.02 mol/L HI, the $H_3O^+(aq)$ (hydronium ion) concentration is approximately

A) 10^{-13} mol/L

B) 10^{-12} mol/L

*C) 10^{-2} mol/L

D) 10^{-1} mol/L

7) What is the hydroxide ion concentration, $[OH^-(aq)]$, in a 0.2 mol/L nitric acid solution?

A) 2×10^{-14} mol/L

*B) 5×10^{-14} mol/L

C) 5×10^{-7} mol/L

D) 2×10^{-1} mol/L

13A.6.

8) 150 mL of 0.2 mol/L NaOH (sodium hydroxide) is added to 50 mL of 0.1 mol/L H_2SO_4 (sulfuric acid). Assuming complete dissociation of both the acid and the base and complete neutralization, the resultant concentration of H^+ is

A) 8×10^{-14} mol/L
*B) 1×10^{-13} mol/L
C) 1×10^{-7} mol/L
D) 1×10^{-1} mol/L

9) 11.4 mL of 0.70 mol/L NaOH (sodium hydroxide) were used to titrate 0.50 g of a solid acid with relative molecular mass 126. What is the number of moles of hydrogen ion per mole of the acid?

A) 1
*B) 2
C) 3
D) 4

10) What volume of 4.00×10^{-2} mol/L HCl (hydrochloric acid) solution is needed to completely react with 2.00×10^{2} mL of 1.00×10^{-2} mol/L LiOH (lithium hydroxide) solution?

*A) 50.0 mL
B) 1.00×10^{2} mL
C) 2.00×10^{2} mL
D) 5.00×10^{2} mL

11) What volume of 0.200 mol/L $Ca(OH)_2$ (calcium hydroxide) solution is required to neutralize 25.0 mL of 0.100 mol/L HCl (hydrochloric acid)?

*A) 6.25 mL
B) 12.5 mL
C) 25.0 mL
D) 50.0 mL

12) What volume of 0.16 mol/L HCl(aq) (hydrochloric acid) is required to completely neutralize 40 mL of 0.20 mol/L NaOH(aq) (sodium hydroxide)?

A) 30 mL
B) 40 mL
*C) 50 mL
D) 60 mL

13) The best choice of indicators for the titration of a weak acid (sample) with a strong base (titrant) is

A) bromothymol blue
*B) phenolthalein
C) litmus
D) methyl red

14) If the concentration of acid in a solution is increased, the

*A) $[H_3O^+]$ increases and pH decreases.

B) $[H_3O^+]$ decreases and pH increases.

C) $[H_3O^+]$ and pH both increase.

D) $[H_3O^+]$ and pH both decrease.

15) A solution has a pH of 5. The $[H^+]$ and $[OH^-]$ in the solution are respectively

*A) 10^{-5} and 10^{-9} mol/L
B) 10^{-9} and 10^{-5} mol/L
C) 10^{+5} and 10^{+9} mol/L
D) 10^{+9} and 10^{+9} mol/L

13A.6.

16) What is the concentration of H_3O^+ (hydronium ion) in a solution whose pH is 4?

A) 10^{+4} mol/L
*B) 10^{-4} mol/L
C) 10^{-10} mol/L
D) 4 mol/L

17) The pH of 0.001 mol/L HCl (hydrochloric acid) solution is

A) 1
B) 2
*C) 3
D) 4

18) What is the pH of a 0.015 mol/L aqueous solution of HCl (hydrochloric acid)?

A) 0.015
B) 0.085
*C) 1.82
D) 2.18

19) For the acid CH_3COOH (acetic acid), the correct expression for K_a is

A) $\dfrac{[CH_3COOH]}{[H^+][CH_3COO^-]}$

*B) $\dfrac{[H^+][CH_3COO^-]}{[CH_3COOH]}$

C) $\dfrac{[H^+][CH_3COO^-]}{[H_2O][CH_3COOH]}$

D) $[H^+][CH_3COO^-]$

20) Arrange the 0.1 mol/L aqueous solutions of the following acids in order of decreasing acidity.

I)	Hydrofluoric acid, HF	$K_a = 3.53 \times 10^{-4}$
II)	Hydrocyanic acid, HCN	$K_a = 4.93 \times 10^{-10}$
III)	Nitrous acid, HNO_2	$K_a = 4.60 \times 10^{-4}$
IV)	Hypobromous acid, HBrO	$K_a = 2.06 \times 10^{-9}$

A) II, IV, I, III
B) II, I, IV, III
*C) III, I, IV , II
D) III, IV, I, II

21) Which of the following, dissolved in water at 298K, is the strongest acid?

A)	HClO	$K_a = 3.0 \times 10^{-8}$
B)	$HClO_2$	$K_a = 1.0 \times 10^{-2}$
C)	$HClO_3$	$K_a = 1.0 \times 10^{2}$
*D)	$HClO_4$	$K_a = 1.0 \times 10^{11}$

22) A 0.20 mol/L solution of formic acid has a pH of 5.00. What is the K_a (ionization constant) of formic acid?

*A) 5.0×10^{-10}
B) 1.0×10^{-5}
C) 5.0×10^{-5}
D) 25

23) In a 0.10 mol/L solution of hydrofluoric acid (HF) the concentration of H^+ is 8.2×10^{-3} mol/L. The ionization constant is

A) 8.2×10^{-6}
B) 1.6×10^{-5}
*C) 6.7×10^{-4}
D) 7.5×10^{-4}

24) Given that K_a for acetic acid is 1.8×10^{-5}, what concentration of acetic acid is 2.0% dissociated?

A) 4.4×10^{-4} mol/L
B) 8.8×10^{-4} mol/L
*C) 0.044 mol/L
D) 22 mol/L

25) The acid-base indicator methyl red, may be represented by HMr, and its ionization equilibrium may be represented by the equation

$$HMr = H^+ + Mr^-$$

Methyl red exhibits the following colours in solution at given hydrogen ion concentrations:

COLOUR	red	red	red	orange	yellow	yellow	yellow	yellow
$[H^+]$	10^{-1}	10^{-2}	10^{-3}	10^{-4}	10^{-5}	10^{-6}	10^{-7}	10^{-8}

The equilibrium constant for the indicator is approximately

A) 1.0×10^{-8} mol/L
B) 1.0×10^{-5} mol/L
*C) 1.0×10^{-4} mol/L
D) 1.0×10^{-3} mol/L

26) A Brønsted-Lowry base is defined as

A) a hydroxide ion
B) a proton donor
C) an electron pair acceptor
*D) a proton acceptor

27) Water can act as either an acid or a base. The equation which represents water reacting as an acid is

*A) $H_2O(\ell) + NH_3(g) = OH^-(aq) + NH_4^+(aq)$

B) $H_2O(\ell) = H_2(g) + \frac{1}{2}O_2(g)$

C) $H_2O(\ell) + HCl = H_3O^+(aq) + Cl^-(aq)$

D) $H_2O(\ell) + C(s) = CO(g) + H_2(g)$

28) When a solution of sodium hydrogen carbonate ($NaHCO_3$) reacts with ammonia (NH_3), the Brønsted-Lowry acid is

A) NH_3

*B) HCO_3^-

C) CO_3^{2-}

D) H_2O

29) In which of the following reactions is the hydrogen sulfate ion (HSO_4^-) acting as a base?

*A) $HSO_4^- + H_3O^+ \rightarrow H_2SO_4 + H_2O$

B) $HSO_4^- + H_2O \rightarrow SO_4^{2-} + H_3O^+$

C) $HSO_4^- + NH_3 \rightarrow NH_4^+ + SO_4^{2-}$

D) $HSO_4^- + 2H^+ + 2e^- \rightarrow HSO_3^- + H_2O$

13A.6.

30) The hydrogen ion concentration in an aqueous solution whose pH is 4.70 is

*A) 2.0×10^{-5} mol/L
B) 5.0×10^{-4} mol/L
C) 4.7 mol/L
D) 9.3 mol/L

31) A 0.050 mol/L HCl (hydrochloric acid) solution has a pH of

*A) 1.3
B) 2.0
C) 2.3
D) 5.0

32) 11.4 mL of 0.70 mol/L NaOH solution was required to completely neutralize 0.50 g of a solid acid of molar mass 126 g/mol. The number of moles of hydrogen ion per mole of the solid acid is

A) 1
*B) 2
C) 3
D) 4

33) Which one of the following is an Arrhenius base?

*A) CsOH
B) NaCl
C) NH_3
D) HCN

13A.6.

34) Which one of the following statements about the ion product (K_w) in water is true?

*A) K_w is equal to 1×10^{-14} at 298 K.

B) K_w depends on the acidity of the solution.

C) K_w depends on the atmospheric pressure.

D) K_w is independent of temperature.

35) If a 0.10 mol/L solution of a weak acid is diluted to 0.010 mol/L at constant temperature the

A) hydrogen ion concentration will increase

B) pH will decrease

C) ionization constant of the weak acid will decrease

*D) percentage ionization will increase

36) In which of the following reactions does nitric acid (HNO_3) act as a base?

*A) $HNO_3 + H_2SO_4 \rightarrow H_2NO_3^+ + HSO_4^-$

B) $HNO_3 \rightarrow H^+ + NO_3^-$

C) $Cu + 4HNO_3 \rightarrow Cu^{2+} + 2NO_3^- + 2H_2O + 2NO_2$

D) $HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$

37) The pH of a solution in which the hydroxide ion, OH^-, concentration is 2.00×10^{-3} mol/L is

A) 1.70

B) 2.70

C) 11.0

*D) 11.3

38) Consider the following aqueous solutions of acids:

I) 1.0 mol/L H_2SO_4 (K_a very large)

II) 1.0 mol/L HCN ($K_a = 1.0 \times 10^{-10}$)

III) 1.0 mol/L HBr (K_a very large)

IV) 1.0 mol/L CH_3COOH ($K_a = 1.8 \times 10^{-5}$)

11 mL samples of 1.0 mol/L of NaOH (sodium hydroxide) solution are added to 10 mL samples of each of the above acids. A sample of phenolphthalein indicator (red in base, colourless in acid) will turn colourless in

*A) I only
B) I and III only
C) III only
D) II and IV only

39) The acid-base indicator methyl red, may be represented by HMr, and its ionization equilibrium may be represented by the equation

$$HMr = H^+ + Mr^-$$

Methyl red exhibits the following colours in solution at given hydrogen ion concentrations:

COLOUR	red	red	red	orange	yellow	yellow	yellow	yellow
$[H^+]$	10^{-1}	10^{-2}	10^{-3}	10^{-4}	10^{-5}	10^{-6}	10^{-7}	10^{-8}

The yellow colour is due primarily to aqueous

*A) Mr^-
B) H^+
C) HMr
D) $HMr + Mr^-$

40) The acid-base indicator methyl red, may be represented by HMr, and its ionization equilibrium may be represented by the equation

$$HMr = H^+ + Mr^-$$

Methyl red exhibits the following colours in solution at given hydrogen ion concentrations:

COLOUR	red	red	red	orange	yellow	yellow	yellow	yellow
$[H^+]$	10^{-1}	10^{-2}	10^{-3}	10^{-4}	10^{-5}	10^{-6}	10^{-7}	10^{-8}

The equilibrium constant for this indicator is

A) 10^{-14}

B) 10^{-8}

C) 10^{-5}

*D) 10^{-4}

41) The acid-base indicator methyl red, may be represented by HMr, and its ionization equilibrium may be represented by the equation

$$HMr = H^+ + Mr^-$$

Methyl red exhibits the following colours in solution at given hydrogen ion concentrations:

COLOUR	red	red	red	orange	yellow	yellow	yellow	yellow
$[H^+]$	10^{-1}	10^{-2}	10^{-3}	10^{-4}	10^{-5}	10^{-6}	10^{-7}	10^{-8}

Methyl red would best be used when titrating

A) a weak acid and a strong base

*B) a strong acid and a weak base

C) a strong acid and a strong base

D) a weak acid and a weak base

13A.6.

42) Which one of the following species can be a Brønsted-Lowry base, but not an Arrhenius base?

*A) Cl^-

B) NaOH

C) $Ca(OH)_2$

D) Ba^{2+}

43) The pH of a 0.020 mol/L aqueous solution of HClO, hypochlorous acid, ($K_a = 4.0 \times 10^{-8}$) is

A) 1.70

B) 3.70

*C) 4.55

D) 7.40

44) The pH of a solution is 5. The hydrogen ion and hydroxide ion concentrations (mol/L) are, respectively,

A) $[H^+] = 10^{-9}$, $[OH^-] = 10^{-5}$

B) $[H^+] = 10^{-5}$, $[OH^-] = 10^{-5}$

*C) $[H^+] = 10^{-5}$, $[OH^-] = 10^{-9}$

D) $[H^+] = 10^{-7}$, $[OH^-] = 10^{-7}$

13A.6.

45) CH_3COOH (acetic acid) ionizes in water according to the equation:

$$CH_3COOH(aq) = H^+(aq) + CH_3COO^-(aq)$$

If CH_3COONa (sodium acetate) is added to the acid, which one of the following will occur?

A) the hydrogen ion concentration will increase
*B) the hydrogen ion concentration will decrease
C) a precipitate will form
D) the hydrogen ion concentration will be unchanged

46) Which of the following best describes what occurs when ammonia gas (NH_3) is bubbled into water?

*A) NH_3 acts as a base towards water
B) NH_3 acts as an acid towards water
C) NH_3 does not dissolve in water
D) no acid-base reaction occurs

47) What volume of 0.0400 mol/L H_2SO_4 (sulfuric acid) is required to completely neutralize 200 mL of 0.0200 mol/L $Al(OH)_3$ (aluminum hydroxide) solution?

A) 50.0 mL
B) 100 mL
*C) 150 mL
D) 300 mL

13A.6.

48) The pH of an aqueous solution of sodium hydrogen carbonate ($NaHCO_3$) is

*A) greater than 7 because the hydrogen carbonate ions accept protons

B) greater than 7 because the hydrogen carbonate ions form carbonate ions by reaction with the water

C) 7 because sodium hydrogen carbonate is a salt

D) less than 7 because the hydrogen carbonate ions dissociate

49) The equation for the ionization of HA is:

$$HA(aq) = H^+(aq) + A^-(aq)$$

The expression for the ionization constant K_a for the acid HA is

A) $\frac{[HA]}{[H^+]\ [A^-]}$

B) $[H^+]\ [A^-]$

C) $\frac{1}{[HA]}$

*D) $\frac{[H^+]\ [A^-]}{[HA]}$

50) According to both the Arrhenius and Brønsted-Lowry definitions, an acid is a compound

A) which contains hydrogen or hydrogen ions

B) whose aqueous solution has a pH greater than 7

C) which can react with hydroxide ions

*D) which ionizes in solution to produce protons

13A.6.

51) Given the following dissociation constants:

$$X : K_b = 1.8 \times 10^{-5}$$
$$Y : K_b = 1.6 \times 10^{-9}$$
$$Z : K_b = 3.8 \times 10^{-10}$$

the arrangement of the substances X, Y and Z in order of increasing base strength is

A) X Y Z
*B) Z Y X
C) Y Z X
D) Y X Z

52) 10 mL of 0.10 mol/L HCl(aq) (hydrochloric acid) requires 35 mL of NaOH(aq) (sodium hydroxide) for complete neutralization. The concentration of the NaOH(aq) is

A) 0.020 mol/L
*B) 0.029 mol/L
C) 0.035 mol/L
D) 0.058 mol/L

53) The hydrogen ion concentration, $[H^+(aq)]$, of a solution is found to be 0.0010 mol/L. The pH of the solution is

A) 1.0
B) 10
*C) 3.0
D) 14

54) The pOH of 0.00010 mol/L nitric acid (HNO_3) solution is

A) 1.0×10^{-10}
B) 1.0×10^{-4}
*C) 10
D) 4.0

55) Consider the following acid ionization constants, K_a:

$CH_3COOH \quad K_a = 1.8 \times 10^{-5}$

$NH_4^+ \quad K_a = 5.6 \times 10^{-8}$

$HIO_3 \quad K_a = 1.9 \times 10^{-1}$

The acids listed in order of increasing acid strength are

A) CH_3COOH NH_4^+ HIO_3

B) HIO_3 CH_3COOH NH_4^+

*C) NH_4^+ CH_3COOH HIO_3

D) NH_4^+ HIO_3 CH_3COOH

56) For the equilibrium:

$$HSO_4^- + HPO_4^{2-} = SO_4^{2-} + H_2PO_4^-$$

a conjugate Brønsted-Lowry acid-base pair is

A) HSO_4^- (base) and SO_4^{2-} (acid)

*B) HPO_4^{2-} (base) and $H_2PO_4^-$ (acid)

C) HSO_4^- (acid) and HPO_4^{2-} (base)

D) HPO_4^{2-} (acid) and $H_2PO_4^-$ (base)

57) The pH of a 0.20 mol/L solution of formic acid (CHOOH) is 5.0. The K_a of formic acid is

*A) 5.0×10^{-10}

B) 5.0×10^{-5}

C) 0.20

D) 25

58) In basic solutions, $[H^+(aq)]$ is

A) $> [OH^-(aq)]$

*B) $< 10^{-7}$ mol/L

C) 10^{-7} mol/L

D) $> 10^{-7}$ mol/L

59) If a small amount of concentrated hydrochloric acid (HCl(aq)) were added to a solution of a weak acid, oxalic acid ($H_2C_2O_4$), which one of the following would occur?

A) $[H_2C_2O_4(aq)]$ would decrease and $[H^+(aq)]$ would increase

B) $[HC_2O_4^-(aq)]$ and $[H^+(aq)]$ would both increase

C) $[H_2C_2O_4(aq)]$ and $[HC_2O_4^-(aq)]$ would both decrease

*D) $[H_2C_2O_4(aq)]$ would increase and $[HC_2O_4^-(aq)]$ would decrease

60) Which of the following aqueous solutions has the highest pH?

A) 0.1 mol/L acetic acid (CH_3COOH)

B) 0.1 mol/L ammonium hydroxide (NH_4OH)

C) 0.1 mol/L sodium acetate (CH_3COONa)

*D) 0.1 mol/L sodium hydroxide ($NaOH$)

61) HI (hydroiodic acid) is a very strong acid. In an aqueous solution of 0.01 mol/L HI, the hydronium ion $[H_3O^+(aq)]$, concentration is:

A) 10^{-13} mol/L

B) 10^{-12} mol/L

*C) 10^{-2} mol/L

D) 10^{-1} mol/L

13A.6.

62) For hypochlorous acid (HClO) $K_a = 4.0 \times 10^{-8}$. The $[H_3O^+(aq)]$ of a 0.010 mol/L solution of HClO is

A) 4.0×10^{-10} mol/L
*B) 2.0×10^{-5} mol/L
C) 2.0×10^{-4} mol/L
D) 1.0×10^{-2} mol/L

63) 2.0 g of sodium hydroxide (NaOH) is added to enough water to make 1.0 L of solution. The $[H^+(aq)]$ in the solution is

*A) 2.0×10^{-13} mol/L
B) 5.0×10^{-12} mol/L
C) 2.0×10^{-3} mol/L
D) 5.0×10^{-2} mol/L

64) In a 0.010 mol/L ammonia solution, $(NH_3(aq))$

$$K_b(NH_3) = 1.8 \times 10^{-5}$$

the concentration of hydroxide ion, $[OH^-aq]$ is

A) 1.8×10^{-7} mol/L
*B) 4.2×10^{-4} mol/L
C) 1.3×10^{-3} mol/L
D) 1.8×10^{-3} mol/L

65) The $H^+(aq)$ concentration in a solution of pH = 8 is

*A) 10^{-8} mol/L
B) 10^{-6} mol/L
C) 10^{6} mol/L
D) 10^{8} mol/L

66) A 0.10 mol/L aqueous solution of a weak acid has a hydrogen ion concentration of 0.0010 mol/L. The value of the ionization constant, K_a, for this acid is

A) 1.0×10^{-7}

B) 1.0×10^{-6}

*C) 1.0×10^{-5}

D) 1.0×10^{-4}

67) Which of the following is a necessary condition for a water solution to conduct electricity when a small voltage is applied between conducting electrodes in the solution?

*A) The solute dissolved in the water must exist as ions

B) The solute dissolved in the water must be one formed by the sharing of electrons

C) A chemical change to the solute must not take place during the process of conduction

D) The solute dissolved in the water must be a compound in which all of the constituent elements are non-metals

68) An aqueous solution of KCl (potassium chloride) is

A) acidic because HCl is a stronger electrolyte than KOH

B) acidic because KOH is a stronger electrolyte than HCl

*C) nearly neutral because KOH and HCl are both strong electrolytes

D) nearly neutral because KOH and HCl are both weak electrolytes

69) 10 mL of a 0.20 mol/L HCl (hydrochloric acid) solution completely neutralized 25 mL of a NaOH (sodium hydroxide) solution. The concentration of the NaOH solution was

A) 0.050 mol/L
*B) 0.080 mol/L
C) 0.50 mol/L
D) 0.80 mol/L

70) A substance is an Arrhenius acid if it

A) turns phenolphthalein solution red
*B) reacts with water to form hydrogen ions
C) turns litmus paper blue
D) reacts with water to form hydroxide ions

71) The acetate ion (CH_3COO^-) is

A) a triprotic acid because it contains three hydrogen atoms
B) an exceptionally weak base because it shows very little tendency to combine with hydrogen ions
C) neither an acid nor a base
*D) a strong base because it readily combines with hydrogen ions to form acetic acid

72) Methyl red ($K_a = 1.0 \times 10^{-5}$) is a good indicator for the titration of a

A) strong base and weak acid
*B) strong acid and weak base
C) strong acid and strong base
D) weak acid and weak base

13A.6.

73) Which one of the following acids is the weakest?

A) HNO_2 $K_a = 6 \times 10^{-6}$
*B) H_2S $K_a = 1.0 \times 10^{-7}$
C) C_6H_5COOH $K_a = 6.6 \times 10^{-5}$
D) HF $K_a = 6.7 \times 10^{-4}$

74) Consider the following aqueous solutions of acids:

I) 1.0 mol/L H_2SO_4
II) 0.1 mol/L HCN ($K_a = 1.0 \times 10^{-10}$)
III) 0.1 mol/L HBr
IV) 1.0 mol/L CH_3COOH ($K_a = 1.8 \times 10^{-5}$)

Which of the conjugate bases of these acids has the **greatest tendency** to combine with $H^+(aq)$ ions?

A) HSO_4^-
*B) CN^-
C) Br^-
D) CH_3COO^-

75) Which of the following is correctly named?

A) NH_4 ammonia
B) $Mn(CO_3)_2$ magnesium carbonate
*C) NH_4ClO ammonium hypochlorite
D) Na_2SO_3 sodium sulfate

76) An aqueous solution of hydrochloric acid (HCl(aq)) has a concentration of 0.24 mol/L; the pH of the solution is

A) 0.24
B) 0.38
*C) 0.62
D) 1.74

77) Which one of the following aqueous solutions contains the highest concentration of $OH^-(aq)$ (hydroxide) ions?

A) 1.0 mol/L CH_3COOH (acetic acid; $K_a = 1.8 \times 10^{-5}$)

B) 1.0 mol/L H_2SO_4 (sulfuric acid)

C) 0.10 mol/L HBr (hydrobromic acid)

*D) 0.10 mol/L HCN (hydrocyanic acid; $K_a = 1.0 \times 10^{-10}$)

78) The process in which hydrogen and hydroxyl ions unite in a solution to form water is termed

A) hydrogenation

B) oxidation

C) hydrolysis

*D) neutralization

79) The acid-base indicator methyl red may be represented by HMr. Its ionization equilibrium may be represented by the following equation:

$$HMr = H^+ + Mr^-$$

Methyl red exhibits the following colours in solution at the given hydrogen ion concentrations:

COLOUR	red	red	red	orange	yellow	yellow	yellow	yellow
$[H^+]$	10^{-1}	10^{-2}	10^{-3}	10^{-4}	10^{-5}	10^{-6}	10^{-7}	10^{-8}

The yellow colour is due primarily to the presence of

A) HMr

B) H^+

*C) Mr^-

D) $HMr + Mr^-$

13A.6.

80) 11.2 g of potassium hydroxide (KOH) is dissolved in sufficient water to make 1 L of solution. What is the concentration of KOH in the solution?

A) 0.01 mol/L
B) 0.02 mol/L
C) 0.1 mol/L
*D) 0.2 mol/L

81) Which one of the following aqueous solutions contains the highest mol/L concentration of ions?

*A) 1.0 mol/L H_2SO_4
B) 1.0 mol/L HCN ($K_a = 1.0 \times 10^{-10}$)
C) 1.0 mol/L HBr
D) 1.0 mol/L CH_3COOH ($K_a = 1.8 \times 10^{-5}$)

82) Which one of the following will occur if a 0.10 mol/L solution of a weak acid is diluted to 0.010 mol/L at constant temperature?

A) The hydrogen ion concentration will increase.
B) The pH will decrease.
C) The ionization constant of the acid will increase.
*D) The percentage ionization in the solution will increase.

83) Which of the following statements about a 1.0 mol/L solution of potasssium carbonate K_2CO_3 is correct?

A) It reacts with magnesium metal to produce CO_2
B) It reacts with silver metal to produce CO_2
C) It is a neutral solution (pH = 7)
*D) It is a basic solution (pH > 7)

84) If 0.012 mol of solid sodium hydroxide is added to 1 L of 0.010 mol/L hydrochloric acid solution, the pH of the solution will be

A) 2.7
B) 3.3
*C) 11.3
D) 12.7

85) The correct chemical formula for the compound formed by the reaction between the ammonium ion and carbonate ion is

A) $(NH_3)_2CO_3$
B) NH_3CO_3
C) NH_4CO_3
*D) $(NH_4)_2CO_3$

86) The concentration (mol/L) of hydroxide ion, $[OH^-(aq)]$ in a 1.0×10^{-3} mol/L sodium hydroxide (NaOH) solution is

A) 1.0×10^{-11}
*B) 1.0×10^{-3}
C) 1.0×10^{3}
D) 1.0×10^{11}

87) The pH of a 1.0×10^{-5} mol/L $Ca(OH)_2$ (calcium hydroxide) solution is

A) 4.7
B) 5.0
C) 9.0
* D) 9.3

88) What is $[OH^-]$ for pure water at 25^oC?

A) 0 mol/L
B) 1×10^{-14} mol/L
*C) 1×10^{-7} mol/L
D) 7 mol/L

13A.6.

89) Which of the following substances is an **Arrhenius base?**

A) NH_3

B) CN^-

*C) $Ca(OH)_2$

D) BaO

90) A certain compound is a very poor conductor of electricity but can dissolve many other substances forming solutions which conduct electricity well. Its pH is 7, and it is formed when a base reacts with an acid. The compound is

A) acetic acid
*B) water
C) sodium chloride
D) alcohol

91) The $[OH^-(aq)]$, hydroxide ion, concentration in a 0.0001 mol/L $Ca(OH)_2$ (calcium hydroxide) solution is

A) 1×10^{-10} mol/L
B) 0.00005 mol/L
C) 0.0001 mol/L
*D) 0.0002 mol/L

92) 11.2 g of KOH (potassium hydroxide) is dissolved in sufficient water to make 1 L of solution. The $[H^+(aq)]$, hydrogen ion concentration, in the solution is

*A) 5×10^{-14} mol/L
B) 5×10^{-13} mol/L
C) 1×10^{-7} mol/L
D) 2×10^{-1} mol/L

13A.6.

93) The density of a 40%, by mass, H_3PO_4 (phosphoric acid) solution is 1.47 g/mL. What is the concentration, in mol/L, of the phosphoric acid?

A) 6.0×10^{-3}

B) 6.0×10^{-2}

*C) 6.0

D) 15

94) Consider the following ionization constant information for aqueous oxalic acid ($H_2C_2O_4$) solutions at 298 K.

First Ionization

$$H_2C_2O_4(aq) = H^+(aq) + HC_2O_4^-(aq)$$

$$K_a = \frac{[H^+]\ [HC_2O_4^-]}{[H_2C_2O_4]} = 5.90 \times 10^{-2}$$

Second Ionization

$$HC_2O_4^-(aq) = H^+(aq) + C_2O_4^{2-}(aq)$$

$$K_a = \frac{[H^+]\ [C_2O_4^{2-}]}{[HC_2O_4^-]} = 6.4 \times 10^{-5}$$

The species present in the greatest concentration at equilibrium at 298 K in a 1.0 mol/L solution of oxalic acid is

*A) $H_2C_2O_4(aq)$

B) $HC_2O_4^-(aq)$

C) $C_2O_4^{2-}(aq)$

D) $H^+(aq)$

13A.6.

95) How many moles of HCl (hydrochloric acid) are required to completely neutralize 2×10^2 mL of 2×10^{-3} mol/L $Ca(OH)_2$ (calcium hydroxide) solution?

A) 4×10^{-4}
*B) 8×10^{-4}
C) 2×10^{-3}
D) 4×10^{-2}

96) 20 mL of a solution of HCl(aq) (hydrochloric acid) is completely neutralized by 45 mL of 0.80 mol/L KOH(aq) (potassium hydroxide). The concentration of the HCl(aq) is

A) 0.90 mol/L
B) 1.1 mol/L
C) 1.4 mol/L
*D) 1.8 mol/L

97) Which one of the following aqueous solutions contains the highest concentration of $H^+(aq)$ (hydrogen) ions?

A) 1.0 mol/L CH_3COOH (acetic acid; $K_a = 1.8 \times 10^{-5}$)
*B) 0.1 mol/L H_2SO_4 (sulphuric acid)
C) 0.1 mol/L HBr (hydrobromic acid)
D) 1.0 mol/L HCN (hydrocyanic acid; $K_a = 1.0 \times 10^{-10}$)

98) The degree of ionization of a solution refers to

A) the temperature at which the solution ionizes
*B) the fraction of molecules ionized in the solution
C) the quantity of free electrons present in the solution
D) the fraction of unionized molecules in the solution

13A.6.

99) A solution of a salt in water will usually be acidic if the salt is a salt of a

A) weak acid and a strong base
*B) strong acid and a weak base
C) weak acid and a weak base
D) strong acid and strong base

100) What volume of 6.0 mol/L HCl (hydrochloric acid) would be required to prepare 900 mL of 0.20 mol/L HCl solution?

A) 1.3 mL
B) 3.0 mL
*C) 30 mL
D) 7.5×10^2 mL

101) What volume of 0.20 mol/L H_2SO_4(aq) (sulfuric acid) is required to completely neutralize 1.0×10^2 mL of 0.20 mol/L NaOH(aq) (sodium hydroxide)?

A) 20 mL
B) 25 mL
*C) 50 mL
D) 1.0×10^2 mL

102) 25 mL of KOH (potassium hydroxide) solution completely neutralizes 2.0×10^2 mL of HCl (hydrochloric acid). The HCl is a 0.50 mol/L solution. The concentration, in mol/L, of the KOH solution is

A) 0.50
B) 2.0
C) 8.0
*D) 4.0

13A.6.

103) The $[H^+]$ of a solution whose pH is 4.7 is

A) 1.0×10^{-5} mol/L
*B) 2.0×10^{-5} mol/L
C) 2.0×10^{-3} mol/L
D) $1.0 \times 10^{4.7}$ mol/L

104) Consider the following aqueous solutions:

I) 1.0 mol/L CH_3COOH (acetic acid; $K_a = 1.8 \times 10^{-5}$)
II) 1.0 mol/L HCl (hydrochloric acid)
III) 1.0 mol/L HF (hydrofluoric acid; $K_a = 3.53 \times 10^{-4}$)
IV) 1.0 mol/L HCN (hydrocyanic acid; $K_a = 1.0 \times 10^{-10}$)

10 mL samples of the acids are added, respectively to four 10 mL samples of 1.0 mol/L NaOH (sodium hydroxide) solution. The pH values of the four resultant solutions will be

A) 7.0
B) > 7.0
C) < 7.0
*D) all different

105)

0.01 mol/L Na_2SO_4
0.01 mol/L NH_4Cl
0.01 mol/L H_2SO_4
0.01 mol/L HCl

The arrangement of the four solutions in order of increasing pH is

A) NH_4Cl, H_2SO_4, Na_2SO_4, HCl
B) HCl, H_2SO_4, Na_2SO_4, NH_4Cl
*C) H_2SO_4, HCl, NH_4Cl, Na_2SO_4
D) NH_4Cl, Na_2SO_4, HCl, H_2SO_4

106) 2.0 L of 0.5 mol/L solutions of

Acetic acid, CH_3COOH, $K_a = 1.5 \times 10^{-5}$
Hydrochloric acid, HCl
Hydrofluoric acid, HF, $K_a = 3.53 \times 10^{-4}$
Sulphuric acid, H_2SO_4

are added to individual samples of magnesium metal each of mass 1.0 g. The total volume of hydrogen gas produced is

A) greatest with the sulphuric acid
B) greatest with the acetic acid
*C) about the same with all four acids
D) different with all four acids

107) With reference to the reaction:

$$HClO_4 + H_2SO_4 = H_3SO_4^+ + ClO_4^-$$

a conjugate Brønsted-Lowry acid-base pair is

A) $HClO_4$ and H_2SO_4

B) H_2SO_4 and ClO_4^-

*C) $HClO_4$ and ClO_4^-

D) H_3SO_4 and ClO_4^-

108) Which one of the following statements is correct?

*A) Bases turn phenolphthalein pink and litmus blue.
B) Bases turn phenolphthalein pink and litmus red.
C) Bases turn phenolphthalein colourless and litmus red.
D) Bases turn phenolphthalein colourless and litmus blue.

13A.6.

109) $[H^+]$ for a 0.010 mol/L HOCl (hypochlorous acid) solution is

($K_a = 4.0 \times 10^{-8}$ mol/L)

A) 4.0×10^{-10} mol/L
B) 2.0×10^{-8} mol/L
C) 4.0×10^{-6} mol/L
*D) 2.0×10^{-5} mol/L

110) A solution of HNO_3 (nitric acid) is 40% solute by mass and has a density of 1.3 g/cm^3. What is the concentration (in mol/L) of this solution?

A) 1.0
B) 5.2
*C) 8.3
D) 9.4

111) The process of removing hydrogen ions from a solution by adding an equal number of hydroxyl ions is called

A) ionization
B) hydrolysis
*C) neutralization
D) precipitation

112) An acid-base indicator is a substance

A) which changes colour when the pH of the test solution is 7
B) in which the positive and negative ions have a different colour
C) which is used to detect whether the test solution contains an excess of positive or of negative ions
*D) which changes colour when passing through a definite pH range, characteristic of that indicator

13A.6.

113) 0.010 mol of NaOH(s) (sodium hydroxide) is added to 1.0 L of 0.10 mol/L HCl (hydrochloric acid). After equilibrium is established, the concentration of H^+- (aq) and OH^-(aq) respectively are

	$[H^+]$	$[OH^-]$
A)	1.0×10^{-7} mol/L	1.0×10^{-7} mol/L
*B)	9.0×10^{-2} mol/L	1.1×10^{-13} mol/L
C)	9.0×10^{-2} mol/L	1.1×10^{-12} mol/L
B)	1.0×10^{-1} mol/L	1.0×10^{-13} mol/L

114) The best indicator to use for the titration of a strong base with a weak acid is

*A) phenolphthalein
B) bromothymol blue
C) methyl red
D) litmus

115) What is the concentration of H_3O^+ (hydronium ion) in mol/L in a solution whose pH is 4?

A) 10^{+4}
*B) 10^{-4}
C) 10^{-10}
D) 4

116) What is the concentration of H_3O^+ (hydronium ion) in mol/L in 0.10 mol/L NaOH (sodium hydroxide) solution?

A) 1.0×10^{-15}
*B) 1.0×10^{-13}
C) 1.0×10^{-1}
D) $1.0 \times 10^{0.1}$

13A.6.

117) 11.2 g of KOH (potassium hydroxide) is dissolved in sufficient water to make 1 L of solution. The $[H^+]$, hydrogen ion concentration, in the solution is

A) 2×10^{-14} mol/L
*B) 5×10^{-14} mol/L
C) 5×10^{-7} mol/L
D) 2×10^{-1} mol/L

118) 10 mL of 0.050 mol/L sodium hydroxide solution is added to 10 mL of 0.050 mol/L acetic acid solution; the pH of the resultant solution is in the range

A) 0 - 3.5
B) 3.5 - 6.9
C) 6.9 - 7.1
*D) 7.1 - 14

119) What mass of HNO_3 (nitric acid) is present in 50 mL of a 1.5 mol/L solution of HNO_3?

A) 2.1 g
*B) 4.7 g
C) 7.5 g
D) 18.9 g

120) The pH of an aqueous solution is 10.70; the hydrogen ion concentration is closest to

A) 5.0×10^{-7} mol/L
*B) 5.0×10^{-10} mol/L
C) 5.0×10^{7} mol/L
D) 2.0×10^{7} mol/L

13A.6.

121) What mass of H_3PO_4 (phosphoric acid) is needed to react completely with 30 g of Ca (calcium)?

$$3Ca + 2H_3PO_4 \rightarrow Ca_3(PO_4)_2 + 3H_2$$

A) 20 g
*B) 49 g
C) 74 g
D) 98 g

122) Acetic acid has an ionization constant of 1.8×10^{-5} at a particular temperature. The hydrogen ion concentration in a 0.20 mol/L solution of acetic acid is

A) 1.8×10^{-6} mol/L
B) 3.6×10^{-6} mol/L
C) 9.0×10^{-5} mol/L
*D) 1.9×10^{-3} mol/L

123) When 0.1 mol/L sodium hydroxide (NaOH) solution is titrated into 0.1 mol/L acetic acid (CH_3COOH) solution using bromothymol blue indicator,the pH at the

A) end point and the stoichiometric equivalence point are the same
B) end point is greater than seven
*C) equivalence point is greater than seven
D) equivalence point is seven

13A.6.

124) If the acid concentration of a solution is increased the

*A) $[H_3O^+]$ increases and pH decreases

B) $[H_3O^+]$ decreases and pH increases

C) $[H_3O^+]$ and pH both increase

D) $[H_3O^+]$ and pH both decrease

125) The aqueous solution of an acid in which the concentration of hydrogen ion is nearest to 4×10^{-3} mol/L is

A) 1.0 mol/L H_2SO_4

B) 0.1 mol/L HCN ($K_a = 1.0 \times 10^{-10}$)

C) 0.1 mol/L HBr

*D) 1.0 mol/L CH_3COOH ($K_a = 1.8 \times 10^{-5}$)

126) 25 mL of a 2.0 mol/L solution of nitric acid (HNO_3) was titrated with a solution of barium hydroxide ($Ba(OH)_2$) until the indicator showed a neutral solution. 100 mL of the $Ba(OH)_2$ solution was required; the concentration of the $Ba(OH)_2$ solution in mol/L is

A) 1.0
B) 2.0
*C) 0.25
D) 0.50

127) The pH of 0.100 mol/L CH_3COOH (acetic acid) solution is

($K_a = 1.80 \times 10^{-5}$)

A) 1.00
*B) 2.87
C) 3.13
D) 13.0

13A.6.

128) The reaction between the hydroxide ion of a base and the hydrogen ion of an acid is referred to as

A) hydrolysis
B) hydration
*C) neutralization
D) hydrogenation

129) The $[H^+(aq)]$, hydrogen ion concentration (mol/L) in an aqueous sodium hydroxide solution (2×10^{-3} mol/L) is

A) 2×10^{-12}
*B) 5×10^{-12}
C) 5×10^{-11}
D) 2×10^{-3}

130) A solution of 0.01 mol/L HCl (hydrochloric acid) is prepared. On the assumption that the HCl is 100% ionized at this concentration, the concentration of OH^- in this solution is

A) 1.0×10^{-2} mol/L
B) 1.0×10^{-4} mol/L
C) 1.0×10^{-10} mol/L
*D) 1.0×10^{-12} mol/L

131) When salts dissociate in solution, they separate into positive and negative ions. An ion which has a negative charge is

A) hydrogen ion
*B) hydroxide ion
C) ammonium ion
D) calcium ion

13A.6.

132) HA is an acid. Which of the following observations best indicates that HA behaves as a very strong acid in water?

A) A solution of HA has a pH lower than 7.
B) A solution of HA turns litmus paper pink.
C) HA is very soluble in water.
*D) A solution of HA conducts electricity very well.

133) Consider the dissociation of an acid HA:

$$HA\ (aq) = H^+(aq) + A^-(aq)$$

The pH of a 0.025 mol/L solution of HA is 3.0. The K_a for HA is

A) 2.5×10^{-8}
B) 2.5×10^{-7}
C) 4.0×10^{-7}
*D) 4.0×10^{-5}

134) 100 mL of a 0.500 mol/L solution of hydrochloric acid (HCl) will be "just" neutralized by

A) 20.0 g of sodium hydroxide (NaOH)
B) 100 mL of a 0.0500 mol/L NaOH solution
*C) 50.0 mL of a 1.00 mol/L potassium hydroxide (KOH) solution
D) 100 mL of a 0.500 mol/L aluminum hydroxide ($Al(OH)_3$) solution

13A.6.

135) Which of the following equations best represents the dissociation of HPO_4^{2-} in water?

A) $HPO_4^{2-} + H_3O^+ = H_2O + H_2PO_4^-$

B) $HPO_4^{2-} + 2H^+ = H_3PO_4$

*C) $HPO_4^{2-} + H_2O = H_3O^+ + PO_4^{3-}$

D) $HPO_4^{2-} + H_2O = OH^- + H_2PO_4^-$

136) What is the pH of a solution in which the $[OH^-]$ is 1.0×10^{-4} mol/L?

A) -10
B) +7
*C) +10
D) -4

137) An aqueous acid or base is said to be strong if it is

A) highly concentrated
B) in a state of equilibrium
C) highly hydrated
*D) highly dissociated

138) Which one of the following statements is correct?

A) A weak acid ionizes completely in water.
*B) A base reacts with an acid to produce water and a salt.
C) A weak acid does not ionize at all in water.
D) An acid ionizes to produce hydroxide ions in solution.

13A.6.

139) The pH of a neutral solution is

A) 10^{-14}
*B) 7
C) 14
D) 10^{7}

140) Which one of the following species is an **Arrhenius acid?**

A) NO_3^-
*B) HCOOH
C) $Ba(OH)_2$
D) Cl^-

141) An aqueous solution of sodium carbonate (Na_2CO_3)

A) has a pH less than 7
B) is neutral
*C) is basic
D) is acidic

142) 0.985 g of an impure solid sample of sodium hydroxide (NaOH) is completely neutralized by 43.2 mL of 0.275 mol/L hydrochloric acid (HCl) solution. The percent, by mass, NaOH in the solid sample was

A) 1.21
B) 2.42
C) 27.9
*D) 48.2

13A.6.

143) What concentration of CH_3COOH (acetic acid) solution has $[H^+] = 1.0 \times 10^{-3}$ mol/L ?

($K_a = 1.8 \times 10^{-5}$ mol/L)

*A) 0.056 mol/L
B) 0.24 mol/L
C) 18 mol/L
D) 56 mol/L

144) If HA is a stronger acid than HB, then

A) $K_a(HB) \times K_a(HB) = 14$
B) $K_a(HB) = K_a(HA)$
*C) $K_a(HB) < K_a(HA)$
D) $K_a(HB) > K_a(HA)$

13A.7.

1) An electrochemical cell involves the following half reactions:

$$Ag^{+}(aq) + e^{-} \rightarrow Ag\ (s) \qquad E^{\circ} = +\ 0.80\ V$$

$$Mn^{2+}(aq) + 2e^{-} \rightarrow Mn\ (s) \qquad E^{\circ} = -\ 1.18\ V$$

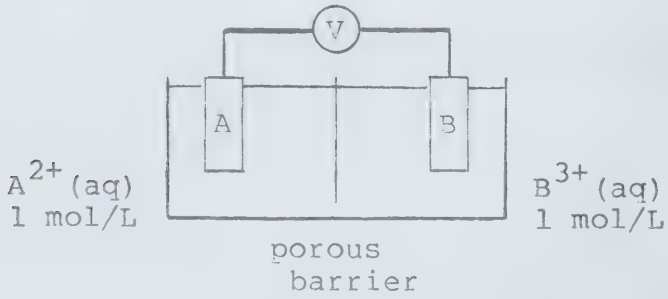
Current is allowed to flow through the cell until the mass of the manganese has changed by 5.49 g. The relative atomic masses of silver and manganese are 108 and 54.9 respectively. The mass of the silver electrode will have changed by

A) 2.75 g
B) 5.49 g
C) 10.0 g
*D) 21.6 g

2) Given the following half cell reactions:

$$A^{2+}(aq) + 2e^{-} \rightarrow A(s) \qquad E^{\circ} = -0.440\ V$$

$$B^{3+}(aq) + 3e^{-} \rightarrow B(s) \qquad E^{\circ} = -0.740\ V$$



With the two electrodes set up in an electrochemical cell at 25°C (as shown in the diagram) the initial cell voltage would be

A) -0.300 V
B) -0.160 V
C) +0.160 V
*D) +0.300 V

3) Given the half reactions:

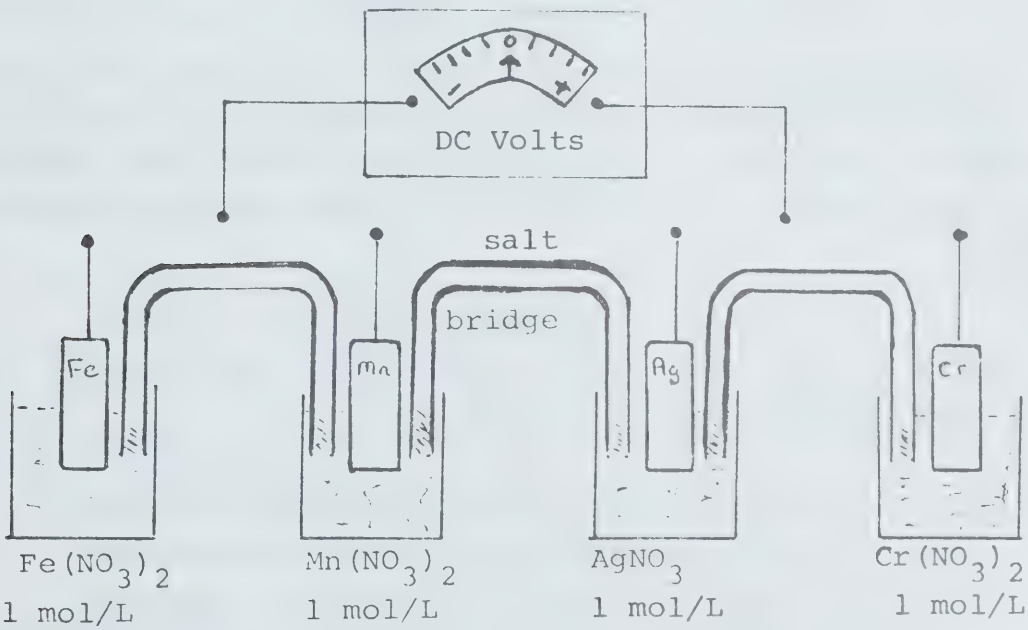
$$Al^{3+} + 3e^- \rightarrow Al(s) \qquad E^o = -1.71\ V$$
$$Cu^+ + e^- \rightarrow Cu(s) \qquad E^o = +0.52\ V$$

Calculate E^o for the reaction:

$$Al^{3+}(aq) + 3Cu(s) \rightarrow Al(s) + 3Cu^+(aq)$$

*A) -2.23 V
B) -1.19 V
C) 1.19 V
D) 2.23 V

4) Consider the cell system shown below in which the wires from the voltmeter may be connected to any pair of electrodes. Assume that no half-reactions other those given can occur.



$Fe^{2+}(aq) + 2e^- \rightarrow Fe\ (s) \qquad E^o = -0.44$ volts
$Mn^{2+}(aq) + 2e^- \rightarrow Mn\ (s) \qquad E^o = -1.18$ volts
$Ag^+(aq) + e^- \rightarrow Ag\ (s) \qquad E^o = +0.80$ volts
$Cr^{3+}(aq) + 3e^- \rightarrow Cr\ (s) \qquad E^o = -0.74$ volts

In order to obtain the highest possible positive voltmeter reading at 25°C the voltmeter leads are connected to the

A) Ag and Fe electrodes
B) Cr and Ag electrodes
*C) Mn and Ag electrodes
D) Mn and Cr electrodes

5) Consider the following half cell reactions:

$$Q^{2+} + 2e^- \rightarrow Q\ (s) \qquad E^\circ = +\ 0.80\ V$$
$$R^{2+} + 2e^- \rightarrow R\ (s) \qquad E^\circ = +\ 0.34\ V$$
$$S^{2+} + 2e^- \rightarrow S\ (s) \qquad E^\circ = -\ 0.25\ V$$
$$T^{2+} + 2e^- \rightarrow T\ (s) \qquad E^\circ = -\ 0.74\ V$$

Which half cell, when combined with a standard hydrogen reference cell under standard conditions, will produce the greatest voltage?

*A) Q
B) R
C) S
D) T

6) An electrochemical cell consists of a chromium electrode in a 1.0 mol/L solution of $Cr(NO_3)_3$ (chromium nitrate), a silver electrode in a 1.0 mol/L solution of $AgNO_3$ (silver nitrate) and a salt bridge. The spontaneous cell reaction is:

$$Cr(s) + 3Ag^+(aq) \rightarrow Cr^{3+}(aq) + 3Ag(s)$$

If a 2.0 mol/L solution of NaCl (sodium chloride) is added to the beaker containing $AgNO_3$,

A) the chromium ion concentration will increase
*B) the reading on the voltmeter will decrease
C) the reading on the voltmeter will increase
D) the reading on the voltmeter will remain unchanged

13A.7.

7) The following half reactions are involved in a voltaic cell:

$$Pb^{2+} + 2e^- \rightarrow Pb \qquad E^\circ = 0.13 \text{ volts}$$
$$Zn \rightarrow Zn^{2+} + 2e^- \qquad E^\circ = -0.76 \text{ volts}$$

The addition of aqueous 0.1 mol/L Na_2SO_4 (sodium sulfate) to the lead half-cell results in the precipitation of $PbSO_4$ (lead sulfate). Which of the following statements is true?

A) The cell voltage will drop to zero.
B) The cell voltage will remain constant.
C) The cell voltage will increase.
*D) The cell voltage will decrease.

8) Consider the following list of half-reactions and the corresponding standard reduction potentials:

$$Na^+ + e^- \rightarrow Na \qquad E^\circ = -2.71 \text{ V}$$
$$I_2 + 2e^- \rightarrow 2I^- \qquad E^\circ = +0.535 \text{ V}$$
$$Fe^{3+} + e^- \rightarrow Fe^{2+} \qquad E^\circ = +0.771 \text{ V}$$
$$Br_2 + 2e^- \rightarrow 2Br^- \qquad E^\circ = +1.07 \text{ V}$$
$$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O \qquad E^\circ = +1.51 \text{ V}$$

Which of the following will oxidize I^- (iodide ion), but not Br^- (bromide ion)?

A) Na
*B) Fe^{3+}
C) MnO_4^- (acidified)
D) Na^+

13A.7.

9) Consider the equation:

$$3N_2O_4 + NaBr + 6NaOH \rightarrow 6NaNO_2 + NaBrO_3 + 3H_2O$$

Which of the following statements is correct?

A) N_2O_4 is the reducing agent, and NaBr is the oxidizing agent.
B) N_2O_4 is the reducing agent, and NaOH is the oxidizing agent.
C) NaBr is the reducing agent, and NaOH is the oxidizing agent.
*D) NaBr is the reducing agent, and N_2O_4 is the oxidizing agent.

10) Which of the statements about the following equation is correct?

$$3Cu(s) + 8HNO_3(aq) \rightarrow 3Cu(NO_3)_2(aq) + 2NO(g) + 4H_2O(\ell)$$

*A) Nitrogen is reduced and copper is oxidized.
B) Nitrogen is reduced and oxygen is oxidized.
C) Nitrogen is oxidized and copper is reduced.
D) Hydrogen is oxidized and oxygen is reduced.

11) Which substance is oxidized in the following reaction?

$$PbO_2(s) + Pb(s) + 4H^+ \rightarrow 2Pb^{2+} + 2H_2O$$

A) PbO_2
*B) Pb
C) H^+
D) Pb^{2+}

13A.7.

12) The oxidation number of Mn (manganese) in $KMnO_4$ (potassium permanganate) is

A) + 6
*B) + 7
C) - 7
D) - 4

13) The oxidation number for bismuth in BiO_3^- is

*A) +5
B) +6
C) +3
D) +7

14) In which of the following does sulfur have an oxidation number of +7?

A) H_2SO_4

B) SO_3

C) HSO_3^-

*D) $H_2S_2O_8$

15) Chromium has an oxidation number of +3 in

A) $Cr_2O_7^{2-}$

B) CrO_2^{2-}

*C) $Cr_2O_4^{2-}$

D) $CrSO_4$

16) According to the following balanced equation:

$$2NaCrO_2 + 2NaOH + NaIO_3 \rightarrow 2Na_2CrO_4 + NaI + H_2O$$

the change in oxidation number for I (iodine) is equivalent to

A) 4 electrons lost
B) 4 electrons gained
C) 6 electrons lost
*D) 6 electrons gained

17) The oxidation number of Cr in $Cr_2O_7^{2-}$ is

*A) +6
B) -6
C) +7
D) +8

18) The oxidation number of sulfur in $Na_2S_2O_3$ is

A) -2
*B) +2
C) -3
D) +3

19) Consider the following equation:

$$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 2Br^-(aq) \rightarrow 2Cr^{3+}(aq) + Br_2(\ell) + 7H_2O(\ell)$$

During this reaction the oxidation number of Cr (chromium) changes from

A) +7 to +3
*B) +6 to +3
C) +7 to 0
D) -2 to +3

20) Which of the substances represented in the following equation undergoes oxidation?

$$4H^+ + 3I_2 + 10NO_3^- \rightarrow 6IO_3^- + 10NO + 2H_2O$$

A) HNO_3
B) HIO_3
*C) I_2
D) NO

21) In the following reaction which element is reduced?

$$2Au(s) + 12H^+(aq) + 3Cl^-(aq) + 3NO_3^-(aq) = 6H_2O(\ell) + 3NOCl(aq) + 2Au^{3+}(aq)$$

A) gold
*B) nitrogen
C) chlorine
D) oxygen

22) In the following reaction which element is reduced?

$$Cr_2O_7^{2-}(aq) + H^+(aq) + Br^-(aq) = Cr^{3+}(aq) + Br_2(\ell) + H_2O(\ell)$$

*A) Cr
B) O
C) H
D) Br

13A.7.

23) Which one of the following half-reactions is completely balanced?

A) $IO_3^- + 6H^+ + 5e^- \rightarrow I_2 + 3H_2O$

B) $SO_4^{2-} + 8H^+ + 8e^- \rightarrow H_2S + 3H_2O$

*C) $ClO^- + H_2O + 2e^- \rightarrow Cl^- + 2OH^-$

D) $NO_2^- + H_2O + 2e^- \rightarrow 2H^+ + NO_3^-$

24) Which one of the following half reaction equations is balanced?

A) $SO_4^{2-} + 8H^+ + 6e^- \rightarrow H_2S + 4H_2O$

B) $NO_2^- + H_2O + 2e^- \rightarrow 2H^+ + NO_3^-$

*C) $ClO^- + H_2O + 2e^- \rightarrow Cl^- + 2OH^-$

D) $2NO_3^- + 4H^+ + 2e^- \rightarrow N_2O + 2H_2O$

25) The unbalanced equation for the Thermite reaction is:

$$Al + Fe_3O_4 \rightarrow Fe + Al_2O_3$$

The sum of the coefficients in the balanced equation is

A) 7
B) 10
C) 15
*D) 24

13A.7.

26) Given the following unbalanced equation:

$$KNO_2 + H_2Cr_2O_7 + HCl \rightarrow KNO_3 + CrCl_3 + H_2O$$

In the balanced equation, the coefficient of HCl is

*A) 6
B) 2
C) 8
D) 10

27) Consider the unbalanced equation:

$$Zn + ___ HNO_3 \rightarrow Zn(NO_3)_2 + N_2O + H_2O$$

The coefficient of HNO_3 in the balanced equation with minimum whole-number coefficients is

A) 8
B) 2
*C) 10
D) 4

28) When silver reacts with dilute nitric acid, the unbalanced equation for the reaction is:

$$Ag(s) + H^+(aq) + NO_3^-(aq) = Ag^+(aq) + NO(g) + H_2O(\ell)$$

The coefficient of NO(g) in the balanced equation is

*A) 1
B) 2
C) 3
D) 4

29) During the electrolysis of water, hydrogen and oxygen gases are produced:

$$2H_2O \rightarrow 2H_2 + O_2$$

How many moles of electrons are required to produce four moles of hydrogen gas, H_2?

*A) 8
B) 2
C) 10
D) 4

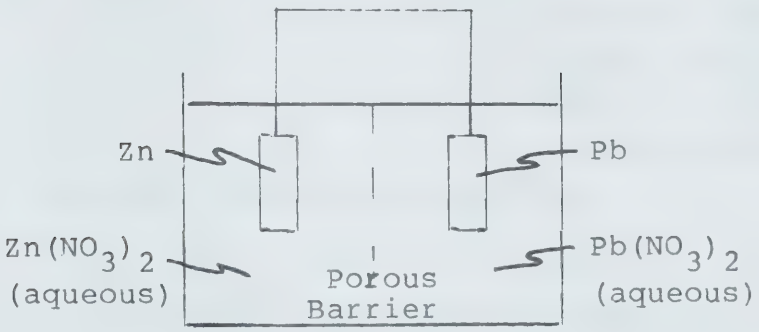
30) Which of the following statements is true for all electrolytic cells?

I) Oxidation takes place at the cathode.
II) Oxidation takes place at the anode.
III) Reduction takes place at the anode.
IV) Reduction takes place at the cathode.

A) I and III
*B) II and IV
C) I and II
D) II and III

31) Consider the following electrochemical cell and standard reduction potentials:

$$Zn^{2+} + 2e^- \rightarrow Zn \qquad E^o = -0.76\ V$$
$$Pb^{2+} + 2e^- \rightarrow Pb \qquad E^o = -0.13\ V$$



The voltage generated by the cell is

A) 0.13 V
*B) 0.63 V
C) 0.89 V
D) 0.76 V

32) Consider the following half cell reactions:

$$Fe^{3+} + 3e^- \rightarrow Fe^o \qquad E^o = -0.44\ V$$
$$Cr^{3+} + 3e^- \rightarrow Cr^o \qquad E^o = -0.74\ V$$
$$Zn^{2+} + 2e^- \rightarrow Zn^o \qquad E^o = -0.76\ V$$
$$Hg^{2+} + 2e^- \rightarrow Hg^o \qquad E^o = +0.78\ V$$
$$Ag^{+} + e^- \rightarrow Ag^o \qquad E^o = +0.80\ V$$

In an electrochemical cell the combination of electrodes (and corresponding electrolytes) that would produce the greatest voltage is

A) Fe - Cr
B) Ag - Hg
*C) Zn - Ag
D) Fe - Ag

13A.7.

33) A metal X reacts with an aqueous solution of $Y(NO_3)_3$ to produce metallic Y and X^+ ions. If 0.3 mol of X are consumed in this reaction, how many moles of Y are produced?

$$X(s) + Y^{3+}(aq) = X^+(aq) + Y(s)$$

*A) 0.1 mol
B) 0.3 mol
C) 0.6 mol
D) 0.9 mol

34) If n coulombs of electricity will deposit 0.1 mol of Ag from a solution of $AgNO_3$ (silver nitrate), how many coulombs will be needed to deposit 0.1 mol of Al from a solution of $Al(NO_3)_3$ (aluminum nitrate)?

A) n

*B) 3n

C) $\frac{n}{3}$

D) $\frac{3 \times 27}{108}$ n

35) Consider the following unbalanced equation:

$$K_2CrO_4 + FeCl_2 + HCl \rightarrow FeCl_3 + CrCl_3 + KCl + H_2O$$

In this reaction

*A) Cr (chromium) is reduced and Fe (iron) is oxidized.
B) Fe (iron) is reduced and Cr (chromium) is oxidized.
C) O (oxygen) is reduced and Cr (chromium) is oxidized.
D) Cr (chromium) is reduced and (H) hydrogen is oxidized.

36) The oxidation number of chlorine in $HClO_4$ (perchloric acid) is

A) -1
B) +1
C) +3
*D) +7

37) Consider the following unbalanced equation:

$$K_2CrO_4 + FeCl_2 + HCl \rightarrow FeCl_3 + CrCl_3 + KCl + H_2O$$

Which of the following substances is the oxidizing agent?

A) $FeCl_3$ (iron(III) chloride)
*B) K_2CrO_4 (potassium chromate)
C) HCl (hydrochloric acid)
D) $FeCl_2$ (iron(II) chloride)

38) The following equation can be balanced using oxidation-reduction half-reactions. The coefficients in the balanced equation are; from left to right

$$H_2O + Cr_2O_7^{2-} + S \rightarrow SO_2 + Cr_2O_3 + OH^-$$

*A) 2, 2, 3, 3, 2, 4
B) 2, 1, 1, 1, 1, 4
C) 2, 2, 3, 2, 2, 4
D) 1, 1, 3, 3, 1, 2

39) Which one of the following groups lists the correct chemical formulae for the hydroxides of aluminum, calcium and potassium?

*A) $Al(OH)_3$ $Ca(OH)_2$ KOH
B) $Al(OH)_2$ $Ca(OH)_2$ KOH
C) Al_3OH Ca_2OH KOH
D) $Al(OH)_3$ CaOH $K(OH)_2$

13A.7.

40) At standard conditions, the species in the list below that is most easily reduced is

A) $Na^+(aq)$

B) $Sn^{2+}(aq)$

*C) $Ag^+(aq)$

D) $Cl^-(aq)$

41) A voltaic cell consists of an electrode of solid lead immersed in a solution of 1.0 mol/L $Pb(NO_3)_2$ (lead nitrate) and an electrode of solid zinc immersed in a solution of 1.0 mol/L $Zn(NO_3)_2$ (zinc nitrate).

$$Pb^{2+} + 2e^- \rightarrow Pb(s) \qquad E^\circ = -0.13\ V$$
$$Zn^{2+} + 2e^- \rightarrow Zn(s) \qquad E^\circ = -0.76\ V$$

The species which will be oxidized is

A) $Zn^{2+}(aq)$

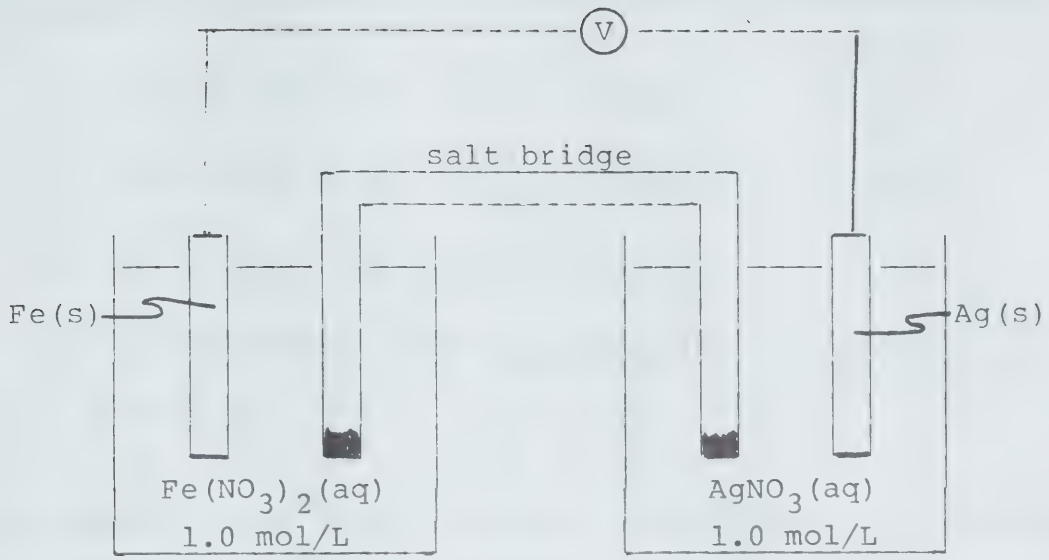
B) $Pb(s)$

C) $Pb^{2+}(aq)$

*D) $Zn(s)$

13A.7.

42)



$$Fe^{2+}(aq) + 2e^- \rightarrow Fe\ (s) \qquad E^\circ = -\ 0.44\ V$$
$$Ag^+(aq) + e^- \rightarrow Ag\ (s) \qquad E^\circ = +\ 0.80\ V$$

Which of the following occurs in the above cell?

A) The Ag electrode decreases in mass.
B) Oxidation occurs at the Ag electrode.
C) The Fe electrode increases in mass.
*D) Oxidation occurs at the Fe electrode.

43) Consider the **unbalanced** equation:

$$H_3AsO_4 + H_2S \rightarrow H_3AsO_3 + S + H_2O$$

In this reaction, As (arsenic) is

A) oxidized from As^{5+} to As^{3+}
B) oxidized from As^{3+} to As^{5+}
*C) reduced from As^{5+} to As^{3+}
D) reduced from As^{3+} to As^{5+}

13A.7.

44) The oxidation number of sulfur in S_8 is

*A) 0
B) +2
C) -2
D) -8

45) Consider the following half cell reactions:

$Q^{2+} + 2e^- \rightarrow Q(s)$ $\quad E^o = + 0.80$ V
$R^{2+} + 2e^- \rightarrow R(s)$ $\quad E^o = + 0.34$ V
$S^{2+} + 2e^- \rightarrow S(s)$ $\quad E^o = - 0.25$ V
$T^{2+} + 2e^- \rightarrow T(s)$ $\quad E^o = - 0.74$ V

Which half cell, when combined with a standard hydrogen reference cell, would produce the greatest voltage difference?

*A) Q
B) R
C) S
D) T

46) $Fe^{3+} + e^- = Fe^{2+}$ $\quad E^o = 0.77$ V
$Fe^{3+} + 3e^- = Fe(s)$ $\quad E^o = -0.036$ V
$Fe^{2+} + 2e^- = Fe(s)$ $\quad E^o = -0.44$ V

The E^o for the reaction:

$$Fe(s) + 2Fe^{3+}(aq) = 3Fe^{2+}(aq)$$

is

A) -0.33 V
B) +0.33 V
C) -1.21 V
*D) +1.21 V

13A.7.

47) Which of the following half-cell reactions is completely balanced?

A) $2IO_3^- + 12H^+ + 5e^- \rightarrow I_2 + 6H_2O$

*B) $ClO^- + H_2O + 2e^- \rightarrow Cl^- + 2OH^-$

C) $NO_2^- + H_2O + 2e^- \rightarrow 2H^+ + NO_3^-$

D) $SO_4^{2-} + 8H^+ + 8e^- \rightarrow H_2S + 4H_2O$

48) The compound most likely to exist at standard conditions is

A) Fe_2O

*B) SiO_2

C) Al_2O

D) PbCl

49) When the following equation is balanced with whole number coefficients, the coefficient of HCl is

$$PbO_2 + HCl \rightarrow PbCl_2 + Cl_2 + H_2O$$

A) 6

B) 2

C) 8

*D) 4

13A.7.

50) Experiments were performed with five metal strips, V, W, X, Y and Z, and their corrresponding 1.0 mol/L nitrate solutions $V(NO_3)_2$, WNO_3, $X(NO_3)_2$, $Y(NO_3)_2$, and $Z(NO_3)_3$.
A strip of metallic Y was inserted into each of the solutions and reactions were observed only in solutions containing X^{2+} and V^{2+} ions.
Metallic W reacted in solutions containing Y^{2+} and Z^{3+} ions but was not tested in the other solutions.
Metallic X did not react in any of the solutions.
A list of the metals in order of decreasing strength as reducing agents (strongest reducing agent first) is

A) V W X Y Z
B) Z Y X W V
*C) W Z Y V X
D) X V Z Y W

51) Consider the following unbalanced equation:

$$Cr_2O_7^{2-}(aq) + H^+(aq) + Br^-(aq) \rightarrow Cr^{3+}(aq) + Br_2(g) + H_2O\ (\ell)$$

When one mole of $Cr_2O_7^{2-}(aq)$ reacts according to this equation, the number of moles of Br_2 produced is

A) 1
B) 2
*C) 3
D) 4

52) An electrochemical cell consists of the following two half-cells:

$$Al^{3+}(aq) + 3e^- \rightarrow Al(s) \qquad E^o = -1.66\ V$$
$$Ag^+(aq) + e^- \rightarrow Ag(s) \qquad E^o = -0.799\ V$$

If three moles of electrons pass through the circuit the number of moles of aluminum which react is

*A) 1
B) 2
C) 3
D) 4

53) Consider the following equation:

$$2H_2SO_4 + 6H_2S \rightarrow S_8 + 8H_2O$$

In this reaction

A) sulphur is oxidized and hydrogen is reduced
B) hydrogen is oxidized and oxygen is reduced
C) sulphur is oxidized and oxygen is reduced
*D) sulphur is both oxidized and reduced

54) During the electrolysis of water, hydrogen and oxygen gas are produced.

$$2H_2O(\ell) \rightarrow 2H_2(g) + O_2(g)$$

The number of moles of electrons required to reduce sufficient water to produce four moles of hydrogen gas (H_2) is

A) 2 moles
B) 4 moles
*C) 8 moles
D) 12 moles

55) In which compound does bromine have the highest oxidation number?

A) KBr
B) KBrO
C) $KBrO_2$
*D) $KBrO_3$

56) Four hypothetical elements M, X, T and Z form the aqueous ions M^{2+}, X^{2+}, T^{2+} and Z^{2+}. Consider the following equations and statements:

$X^{2+} + Z \rightarrow Z^{2+} + X$	reacts appreciably
$X^{2+} + M$	will not react appreciably
$Z^{2+} + T \rightarrow T^{2+} + Z$	reacts appreciably

What is the arrangement of the elements in order of decreasing tendency to lose electrons?

*A)	T	Z	X	M
B)	M	X	Z	T
C)	X	T	Z	M
D)	X	Z	T	M

57) The oxidation state of Cr in $Cr_2O_7^{2-}$ is

A) +5
*B) +6
C) +7
D) +12

58) The formula for aluminum sulphate is $Al_2(SO_4)_3$. The algebraic sum of the oxidation numbers of all of the atoms in the compound is

*A) 0
B) -2
C) +3
D) +6

59) In the formula $H_3X_2O_5$ the oxidation state of the element X is

A) +1
B) +2
*C) +7/2
D) +7

60) Cd (cadmium) may be plated from a 1 mol/L solution of $CdCl_2$ (cadmium chloride) by electrolysis using the object to be plated as the cathode in the electrolytic cell. The equation for the reaction is:

$$CdCl_2(aq) \rightarrow Cd(s) + Cl_2(g)$$

The standard reduction potentials for the two half-reactions are:

$$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq) \qquad E^o = + 1.36 \text{ V}$$

$$Cd^{2+}(aq) + 2e^- \rightarrow Cd(s) \qquad E^o = - 0.400 \text{ V}$$

What is the minimum voltage necessary for the plating reaction to proceed?

A) 0.880 V
B) 0.96 V
*C) 1.76 V
D) 3.52 V

61) Which of the following is the best reducing agent?

A) $F_2(aq)$
B) Mn(s)
*C) Na(s)
D) $Cl^-(aq)$

62) Which of the following half reactions is an example of an oxidation process?

A) $CO_3^{2-} + H_2O \rightarrow HCO_3^- + OH^-$

B) $Zn^{2+} + CO_3^{2-} \rightarrow ZnCO_3$

C) $Cl_2 + 2e^- \rightarrow 2Cl^-$

*D) $Cr^{2+} \rightarrow Cr^{3+} + e^-$

63) Given the following half cell reactions:

$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$ $\quad E^o = + 1.36$ V

$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$ $\quad E^o = + 0.77$ V

Which one of the following reactions may occur spontaneously?

*A) $Cl_2(g) + 2Fe^{2+}(aq) \rightarrow 2Cl^-(aq) + 2Fe^{3+}(aq)$

B) $Cl_2(g) + 2Fe^{3+}(aq) \rightarrow 2Cl^-(aq) + 2Fe^{2+}(aq)$

C) $2Cl^-(aq) + 2Fe^{3+}(aq) \rightarrow Cl_2(g) + 2Fe^{2+}(aq)$

D) $2Cl^-(aq) + 2Fe^{2+}(aq) \rightarrow Cl_2(g) + 2Fe^{3+}(aq)$

64) Consider the metals Q, R, X and Y. The following experimental observations have been made:

$Q(s) + R^{2+}(aq) \rightarrow R(s) + Q^{2+}(aq)$

$X(s) + Y^{2+}(aq) \rightarrow Y(s) + X^{2+}(aq)$

$X(s) + R^{2+}(aq) \rightarrow$ no reaction

The metallic ions, in order of decreasing tendency to acquire electrons, are

A) $Q^{2+}(aq)$ $\quad R^{2+}(aq)$ $\quad X^{2+}(aq)$ $\quad Y^{2+}(aq)$

B) $Y^{2+}(aq)$ $\quad X^{2+}(aq)$ $\quad Q^{2+}(aq)$ $\quad R^{2+}(aq)$

*C) $Y^{2+}(aq)$ $\quad X^{2+}(aq)$ $\quad R^{2+}(aq)$ $\quad Q^{2+}(aq)$

D) $X^{2+}(aq)$ $\quad Y^{2+}(aq)$ $\quad R^{2+}(aq)$ $\quad Q^{2+}(aq)$

13A.7.

65) Consider the following unbalanced equation:

$$N_2H_4 + KMnO_4 + HCl + H_2O \rightarrow MnCl_2 + KCl + NH_2OH$$

When the equation is balanced with the smallest whole-number coefficients the coefficient of $MnCl_2$ will be

A) 1
*B) 2
C) 5
D) 6

66) A electrochemical cell consists of an electrode of solid lead immersed in a solution of 1.0 mol/L $Pb(NO_3)_2$ (lead nitrate) and an electrode of solid zinc immersed in a solution of 1.0 mol/L $Zn(NO_3)_2$(zinc nitrate). The two halves of the cell are separated by a porous barrier; the electrodes are joined by a conducting wire. Within the cell

A) Zn^{2+} ions migrate to the Zn electrode.
*B) Pb^{2+} ions migrate to the Pb electrode.
C) NO_3^- ions migrate from the anode to the cathode.
D) electrons migrate from the anode to the cathode through the porous membrane.

67) Sulphur (S) is treated with nitric acid (HNO_3) to produce sulphuric acid (H_2SO_4). When the equation is balanced, how many moles of nitric acid are required to produce one mole of sulphuric acid?

$$HNO_3(aq) + S(s) \rightarrow NO_2(g) + H_2O(\ell) + H_2SO_4(aq)$$

A) 1
*B) 6
C) 3
D) 4

13A.7.

68) The **oxidizing agent** in the following reaction is

$$I_2(aq) + H_2S(aq) = 2H^+(aq) + 2I^-(aq) + S(s)$$

A) I^-
B) S
C) H_2S
*D) I_2

69) Which of the following is the strongest reducing agent?

A) Ag(s)
*B) Ba(s)
C) $Br_2(aq)$
D) $Fe^{3+}(aq)$

70) The oxidation number of U, uranium, in K_2UO_4 is

A) 0
B) +2
*C) +6
D) +4

71) Consider the **unbalanced** equation

$$FeCl_2 + SnCl_4 \rightarrow FeCl_3 + SnCl_2$$

The sum of the coefficients in the **balanced** equation is

A) 10
B) 8
*C) 6
D) 4

72) When the following equation is balanced the coefficients in order, from left to right, are

$$Cu(s) + HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + NO(g) + H_2O(\ell)$$

A) 3, 2, 3, 2, 1
B) 1, 4, 1, 2, 2
*C) 3, 8, 3, 2, 4
D) 2, 6, 2, 2, 3

73) Consider the following redox reaction:

$$I_2 + HNO_3 \rightarrow HIO_3 + NO + H_2O \quad \text{(unbalanced)}$$

In the balanced equation the coefficient of HIO_3 is

A) 5
*B) 6
C) 3
D) 10

74) Consider the following unbalanced equation:

$$Mg_3N_2 + H_2O \rightarrow Mg(OH)_2 + NH_3$$

When the equation is balanced the sum total of the four coefficients is

A) 6
B) 7
C) 9
*D) 12

75) Which one of the following substances is correctly named?

*A)	$Fe(ClO_2)_3$	iron(III) chlorite
B)	$PbCO_3$	lead(IV) carbonate
C)	$M_3(PO_4)_2$	mercury phosphate
D)	Cu_2SO_4	copper(I) hydrogen sulphite

76) Consider the metals Q, R, X and Y. The following experimental observations have been made

$$Q(s) + R^{2+}(aq) \rightarrow R(s) + Q^{2+}(aq)$$
$$X(s) + Y^{2+}(aq) \rightarrow Y(s) + X^{2+}(aq)$$
$$X(s) + R^{2+}(aq) \rightarrow \text{no reaction}$$

The best reducing agent is

A) Y(s)
*B) Q(s)
C) X(s)
D) R(s)

77) Consider the following half-cell reactions:

$$Fe^{2+}(aq) + 2e^- \rightarrow Fe(s) \qquad E^\circ = -0.44\ V$$
$$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s) \qquad E^\circ = +0.34\ V$$

The relative atomic masses of Fe and Cu are 55.8 and 63.5 respectively.

The two "half cells" are joined by a salt bridge and current is allowed to flow until the mass of the iron electrode has decreased by 2.79 g. During this time the mass of the copper electrode may be expected to have increased by

A) 1.59 g
B) 2.45 g
C) 2.79 g
*D) 3.18 g

13A.7.

78) The tarnish (Ag_2S) on silverware articles can be removed by placing the articles in an aluminum pan and covering them with a warm solution of Na_2CO_3 (sodium carbonate). The following half-reactions are involved:

$$Ag_2S(s) + 2e^- \rightarrow 2Ag(s) + S^{2-}(aq)$$
$$Al(s) + 3OH^-(aq) \rightarrow Al(OH)_3(s) + 3e^-$$

The balanced net ionic equation for the total reaction is

*A) $3Ag_2S(s) + 2Al(s) + 6OH^-(aq) \rightarrow 6Ag(s) + 3S^{2-}(aq) + 2Al(OH)_3(s)$

B) $Ag_2S(s) + Al(s) + 3OH^-(aq) \rightarrow 2Ag^+(aq) + S^{2-}(aq) + Al(OH)_3(s) + e^-$

C) $Ag_2S(s) + Al(OH)_3(s) \rightarrow 2Ag^+(aq) + S^{2-}(aq) + Al(s) + 3OH^-(aq)$

D) $Ag_2S(s) + Al(OH)_3(s) + 3OH^-(aq) \rightarrow Al(s) + 2AgOH(s) + SO_2(g) + 2H_2O$

79) The oxidation number of phosphorus in H_3PO_4 (phosphoric acid) is

A) -5
B) -3
C) +3
*D) +5

13A.7.

80) Experiments were performed with five metal strips, V, W, X, Y and Z, and their corrresponding 1.0 mol/L nitrate solutions, $V(NO_3)_2$, WNO_3, $X(NO_3)_2$, $Y(NO_3)_2$, and $Z(NO_3)_3$.
A strip of metallic Y was inserted into each of the solutions and reactions were observed only in solutions containing X^{2+} and V^{2+} ions.
Metallic W reacted in solutions containing Y^{2+} and Z^{3+} ions.
The metal W was not tested in the other solutions.
Metallic X did not react in any of the solutions.
The metals, listed in order of decreasing activity are

*A) W Z Y V X
B) X V Y Z W
C) Z Y W X V
D) W V Y Z X

13A.8.

1) 196 g of H_3PO_4, (phosphoric acid) is dissolved in enough water to make 1.00 L of solution. The concentration of the phosphoric acid solution is

A) 1.00 mol/L
*B) 2.00 mol/L
C) 3.00 mol/L
D) 0.500 mol/L

2) A solution is prepared by dissolving 5.0 g of solute in 11 g of solvent. The percent by mass of solute in this solution is

A) 28%
*B) 31%
C) 45%
D) 83%

3) What is the simplest formula of a compound containing 40.0% carbon, 6.67% hydrogen and 53.3% oxygen by mass?

*A) CH_2O
B) C_6HO_8
C) CH_8O_3
D) $C_6H_{12}O_6$

4) Compound X contains only Cu (copper) and Cl (chlorine). A 16.4 g sample of compound X contains 10.5 g of Cu. The relative atomic masses of Cu and Cl are 63.5 and 35.5 respectively. The empirical (simplest) formula of the compound X is

*A) CuCl
B) $CuCl_2$
C) Cu_2Cl
D) Cu_2Cl_3

13A.8.

5) A sample of sulfur dioxide, SO_2, is decomposed to produce oxygen gas and sulphur gas. If one volume of $S_8(g)$ is produced, how many volumes of $O_2(g)$ were produced?

*A) 8
B) 2
C) 32
D) 4

6) One Faraday is equivalent to

*A) 9.6489×10^4 C
B) 9.6489×10^4 electrons
C) 9.6489×10^4 V
D) 9.6489×10^4 A

7) An electron beam passes between two oppositely charged plates. The electric field between the two plates is uniform. The electron beam would

A) continue undeflected between the two plates
B) bend in towards the negative plate
*C) bend in towards the positive plate
D) bend upwards out of the field

13A.8.

8) A spectrum was obtained from a mass spectrograph of a sample of hydrogen gas. Due to the presence of a second isotope, D (deuterium), five gaseous ions were produced; H_2^+, HD^+, H^+, D_2^+, D^+, but only four lines appear in the mass spectrum. If the relative atomic mass of deuterium is 2, which two ions form a spectral line at the same position?

A) HD^+ and H^+

B) H_2^+ and HD^+

*C) H_2^+ and D^+

D) HD^+ and D_2^+

9) From the results of the "Gold Foil Experiment", Rutherford concluded that atoms contain very small nuclei because

A) the alpha particles were small

B) a low percentage of alpha particles went straight through the foil

*C) a small number of alpha particles were deflected through large angles

D) gold is such a heavy element

13A.9.

1) The magnitude of a quantum of radiant energy varies directly with

A) speed
B) wavelength
*C) frequency
D) altitude

2) The speed of water waves is 6.2 m/s and their wavelength is 0.50 m. The frequency of the waves is

A) 0.081 Hz
B) 3.1 Hz
C) 6.2 Hz
*D) 12 Hz

3) A photographic plate is darkened more effectively when the light has a lot of energy per photon. Which of the following radiations would produce the darkest plate if each were used in identical procedures?

*A) x-rays
B) ultraviolet
C) infrared
D) radio waves

4) Who postulated that electrons can only make certain energy jumps from level to level?

*A) Bohr
B) Moseley
C) Rutherford
D) Schroedinger

13A.9.

5) How many different values of the quantum number, l, are possible in the 5th energy level of any atom?

A) 25
B) 16
*C) 5
D) 4

6) In modern chemical theory an occupied orbital is pictured to be

A) a spherical or dumbbell-shaped route traced by the electron in its rapid movement.
B) a region in space having a precise shape, which is completely filled by a dense electron cloud.
*C) a region in space in which the probability of finding an electron is high.
D) an elliptical pathway, outside the nucleus, followed by an electron.

7) What is the maximum number of electrons which can be accommodated in the fourth energy (n=4) level of a neutral atom?

A) 16
*B) 32
C) 60
D) 72

13A.9.

8) The atoms of a particular element have the following electron configuration:

$$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$$

One atom of the element has

A) 12 core electrons and 3 valence electrons
B) 11 core electrons and 4 valence electrons
*C) 10 core electrons and 5 valence electrons
D) 8 core electrons and 5 valence electrons

9) The subshell filling pattern of the elements in the sixth row of the periodic table (from left to right) is

A) 6s, 6p, 5d, 4f
*B) 6s, 4f, 5d, 6p
C) 4f, 5d, 6s, 6p
D) 6p, 5d, 4f, 6s

10) The orbital occupancy of the neon ion, Ne^+, is

A) $\underset{1s}{\uparrow}\ \underset{2s}{\uparrow}\ \underset{2p_x}{\uparrow}\ \underset{2p_y}{\uparrow}\ \underset{2p_z}{\uparrow}$

B) $\underset{1s}{\uparrow\downarrow}\ \underset{2s}{\uparrow\downarrow}\ \underset{2p_x}{\uparrow}\ \underset{2p_y}{\uparrow}\ \underset{2p_z}{\uparrow}\ \underset{3s}{\uparrow\downarrow}$

C) $\underset{1s}{\uparrow\downarrow}\ \underset{2s}{\uparrow\downarrow}\ \underset{2p_x}{\uparrow\downarrow}\ \underset{2p_y}{\uparrow\downarrow}\ \underset{2p_z}{\uparrow\downarrow}$

*D) $\underset{1s}{\uparrow\downarrow}\ \underset{2s}{\uparrow\downarrow}\ \underset{2p_x}{\uparrow\downarrow}\ \underset{2p_y}{\uparrow\downarrow}\ \underset{2p_z}{\uparrow}$

13A.9.

11) An element has the following ground state electron configuration for its valence electrons:

$$3s^2\ 3p_x^1\ 3p_y^1 3p_z^1$$

The location of this element, in the Periodic Table is

A) Period 2, Family I
B) Period 3, Family II
C) Period 3, Family III
*D) Period 3, Family V

12) The maximum number of electrons that can occupy the 4th principal energy level of any atom is

A) 8
B) 12
C) 18
*D) 32

13) Which one of the elements whose ground state electron configurations are given below will have the highest first ionization energy?

A) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$

A) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$

*C) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6$

D) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$

14) Ionization energy is the minimum energy required to

A) add an electron to a gaseous atom
*B) remove an electron from a gaseous atom
C) dissolve one mole of an ionic compound
D) vaporize one mole of an ionic compound

13A.9.

15) The least reactive of the inert gas elements in the following list is

A) Xe
B) Ar
C) Kr
*D) He

16) Ionization energies tend to

A) increase with increasing atomic radii
B) decrease with increasing nuclear charge
C) decrease across a period from left to right
*D) increase across a period from left to right

17) The electronic configuration of the sulphide ion, S^{-2}, can be described by

A) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$
*B) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$
C) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$
D) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$

18) The electron configuration of the Mg^{2+} ion (Z = 12) is

A) $1s^2\ 2s^2\ 2p^6\ 3s^2$
B) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$
C) $1s^2\ 2s^2\ 2p^6\ 3s^1$
*D) $1s^2\ 2s^2\ 2p^6$

19) The magnetic quantum number (m_ℓ) describes the

*A) orientation in space of the orbital
B) average distance of the electron from the nucleus
C) shape of the orbital
D) direction of spin of the electron

20) An element of atomic number 84 is represented by the symbol "X". An electron dot diagram for X would be

A) ·X·
B) ·X·
C) X
*D) ·X:

21) The number of half-filled orbitals in the valence shell of an oxygen atom in its ground state is

A) 1
*B) 2
C) 3
D) 0

13A.9.

22) To which families in the periodic table do the following elements belong?

Element Z which forms a Z^{2+} ion
Element Y which forms a Y^{2-} ion

A) Both elements belong to the alkaline earth family.

B) Element Z belongs to the oxygen family and element Y belongs to the alkali metals.

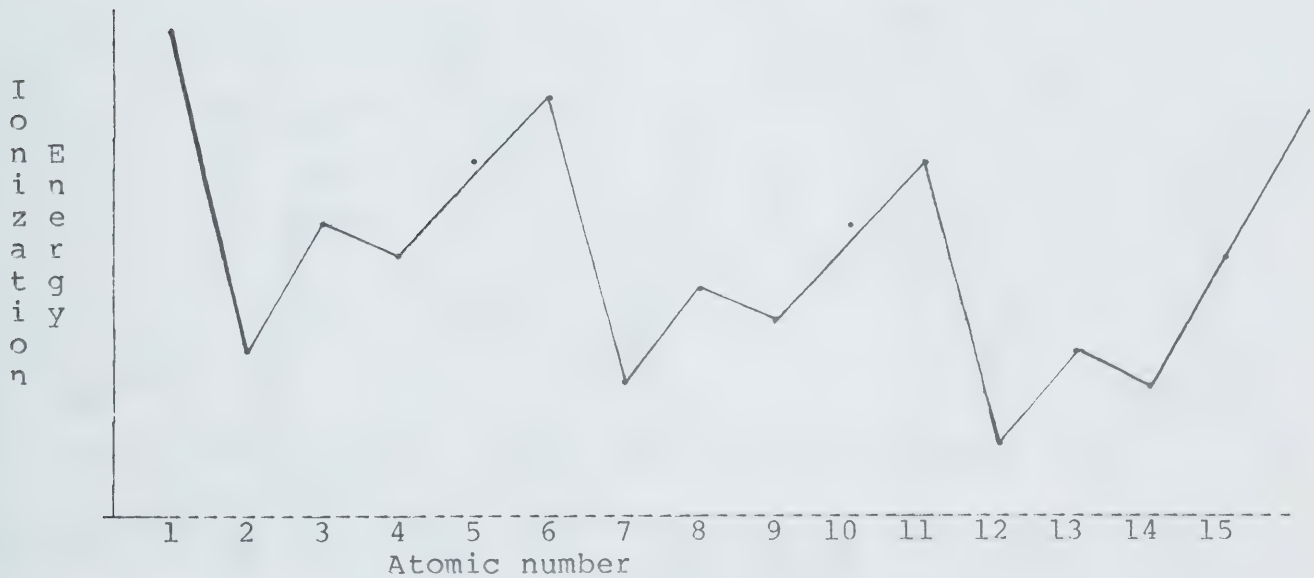
C) Both elements belong to the oxygen family.

*D) Element Z belongs to the alkaline earth family and element Y belongs to the oxygen family.

13A.9.

23) Given below is a list of fifteen fictitious elements and an experimentally produced graph of their first ionization energies.

Atomic Number	Symbol	Name of Element	Atomic Number	Symbol	Name of Element
1	A	aquarius	9	I	inklium
2	B	butyrium	10	J	jothrium
3	C	clutter	11	K	kringlium
4	D	doople	12	L	luciumm
5	E	effressium	13	M	mentium
6	F	floogium	14	N	noxium
7	G	geranium	15	O	oogle
8	H	hoople			



The elements which would be "noble gases" are

*A) aquarius, floogium, and kringlium
B) butyrium, geranium and lucium
C) effressium, jothrium and oogle
D) clutter and hoople

13A.9.

24) Given below is a list of fifteen fictitious elements and an experimentally produced graph of their first ionization energies.

Atomic Number	Symbol	Name of Element	Atomic Number	Symbol	Name of Element
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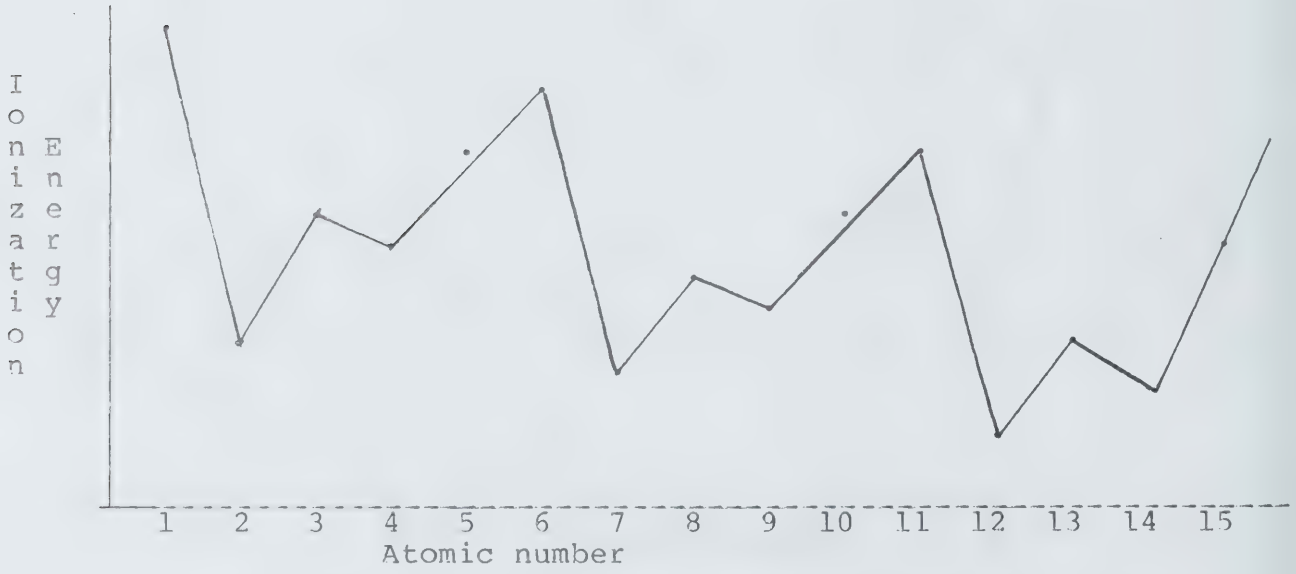
5	E	effressium	13	M	mentium
6	F	floogium	14	N	noxium
7	G	geranium	15	O	oogle
8	H	hoople			



The elements which would be "alkali metals" in this group are

A) aquarius, floogium, and kringlium
*B) butyrium, geranium and lucium
C) effressium, jothrium and oogle
D) clutter and kringlium

13A.9.

25) An ion having 17 protons, 20 neutrons and 18 electrons has a mass number of

A) 35
*B) 37
C) 38
D) 55

26) The correct electronic configuration for $^{107}_{47}Ag$ (silver) is

		4d					5s	5p		
		⎯⎯	⎯⎯	⎯⎯	⎯⎯	⎯⎯	⎯⎯	⎯⎯	⎯⎯	⎯⎯
*A)	$[_{36}Kr]$	↑↓	↑↓	↑↓	↑↓	↑↓	↑			
B)	$[_{36}Kr]$	↑↓	↑↓	↑↓	↑↓	↑	↑	↑		
C)	$[_{36}Kr]$	↑	↑	↑	↑	↑	↑	↑	↑	↑
D)	$[_{36}Kr]$	↑	↑	↑	↑	↑	↑↓	↑	↑	↑

27) The atomic number of an element whose electrons are represented by $1s^2\ 2s^2$ is

A) 6
B) 2
C) 8
*D) 4

13A.9.

28) The following is a list of the usual charges found on the ions of a series of elements:

$$V^{3+},\ W^{2+},\ X^{+},\ Y^{-},\ Z^{2-}$$

Which of the following groups of elements are most likely to be metals?

*A) V, W and X
B) X and Y
C) Y and Z
D) W and Z

29) For the valence shell, the correct representation of carbon in its ground state is

A) $2s^2\ 2p^4$

B) $1s^2\ 2s^2\ 2p^2$

*C) $2s^2\ 2p^2$

D) $3s^2\ 3p^4$

30) Which of the following atoms, in the ground state, would contain an electron with the quantum numbers

$$n = 3,\ \ell = 1,\ m = 1,\ m_s = +\frac{1}{2}$$

A) Na
*B) Cl
C) F
D) Ne

13A.9.

31) The maximum number of electrons which can occupy any one orbital is

A) 6
*B) 2
C) 10
D) 14

32) The formula for an oxide of a Boron family element (X) would be

*A) X_2O_3
B) XO
C) X_3O_2
D) XO_3

33) Which element, in the ground state, is shown with its correct orbital representation?

	$1s^2$	$2s^2$	$2p_x^2$	$2p_y^1$	$2p_z^1$
A) N	↿⇂	↿⇂	↿⇂	↿	‒
B) Li	↿⇂	↿⇂	↿	‒	‒
*C) O	↿⇂	↿⇂	↿⇂	↿	↿
D) Ne	↿⇂	↿⇂	↿	↿	↿

34) Who established the relationship between energy and the
D110 wavelength of electromagnetic radiation, E = h ?

A) Einstein
*B) Planck
C) Bohr
D) De Broglie

13A.9.

35) The ground state electron configuration that will most readily produce a 2+ ion is

A) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$

*B) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2$

C) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^2$

D) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6$

36) If the ionization energy of hydrogen is 1.3×10^3 kJ/mol how much energy is needed to promote an electron from the third to the fifth level of a hydrogen atom?

A) 18 kJ/mol

*B) 92 kJ/mol

C) 1.7×10^2 kJ/mol

D) 3.2×10^2 kJ/mol

37) The valence shell of chlorine may be represented by

*A) $\underline{\uparrow\downarrow}_{3s}\ \underline{\uparrow\downarrow}_{3p}\ \underline{\uparrow\downarrow}_{3p}\ \underline{\uparrow}_{3p}$

B) $\underline{\uparrow}_{3s}\ \underline{\uparrow\downarrow}_{3p}\ \underline{\uparrow\downarrow}_{3p}\ \underline{\uparrow\downarrow}_{3p}$

C) $\underline{\uparrow}_{3s}\ \underline{\uparrow}_{3p}\ \underline{\uparrow}_{3p}\ \underline{\uparrow}_{3p}$

D) $\underline{\uparrow\downarrow}_{3s}\ \underline{\uparrow\downarrow}_{3p}\ \underline{\uparrow}_{3p}\ \underline{\uparrow}_{3p}$

13A.9.

38) For the principal quantum number, n = 1, the permitted values of ℓ, m and m_s respectively are

A) $\ell = 1$ m = -1, 0, +1 $m_s = \mp \frac{1}{2}$

*B) $\ell = 0$ m = 0 $m_s = \mp \frac{1}{2}$

C) $\ell = 1$ m = 0 $m_s = \mp \frac{1}{2}$

D) $\ell = 0$ m = 1 $m_s = \mp \frac{1}{2}$

39) For quantum number $\ell = 2$ the resulting orbital is of type

A) s

B) p

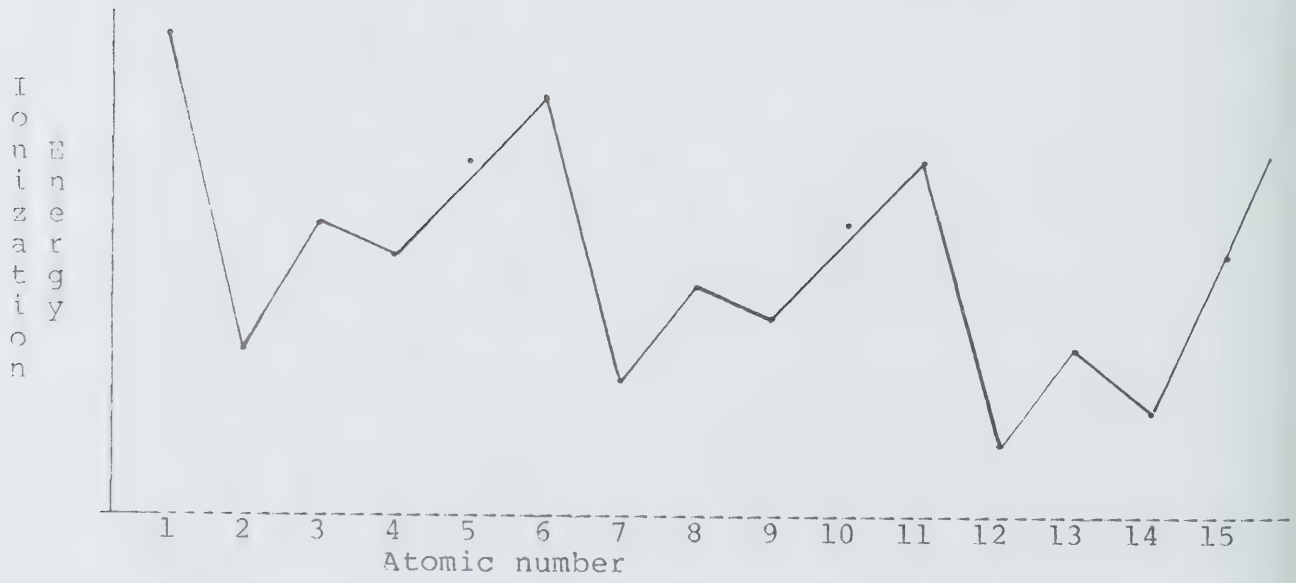
*C) d

D) f

13A.9.

40) Given below is a list of fifteen fictitious elements and an experimentally produced graph of their first ionization energies.

Atomic number	Symbol
1	A
2	B
3	C
4	D
5	E
6	F
7	G
8	H
9	I
10	J
11	K
12	L
13	M
14	N
15	O



In this list of elements, two metallic ions which have the same number of electrons as a neutral F atom are

A) G^{2+} and H^{+}

*B) G^{+} and H^{2+}

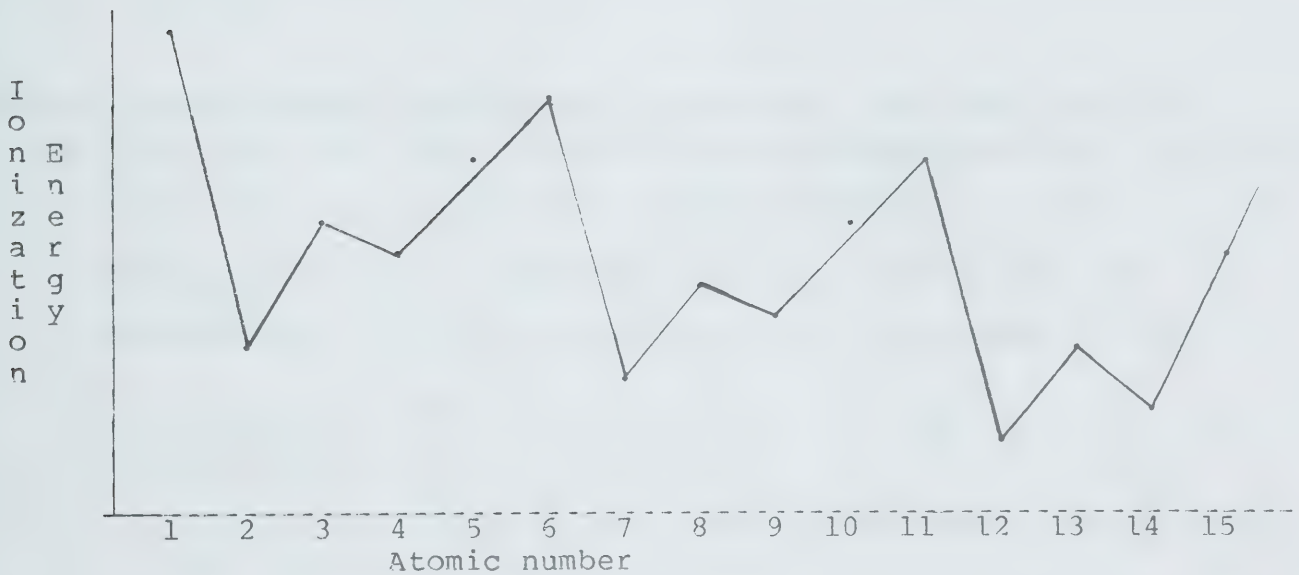
C) D^{2+} and E^{+}

D) D^{2-} and E^{+}

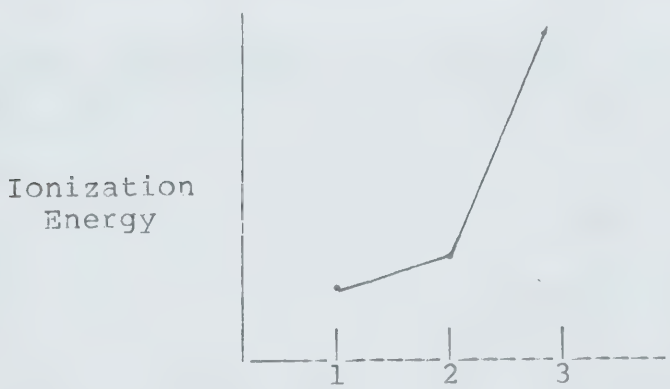
13A.9.

41) Given below is a list of fifteen fictitious elements and an experimentally produced graph of their first ionization energies.

Atomic number	Symbol	Atomic number	Symbol
1	A	9	I
2	B	10	J
3	C	11	K
4	D	12	L
5	E	13	M
6	F	14	N
7	G	15	O
8	H		



Which of the following elements would probably have the plot of successive ionization energies shown below?



A) element B
*B) element C
C) element D
D) element E

13A.9.

42) The Lewis dot diagram for the boron atom with electron configuration $1s^2\ 2s^2\ 2p^1$ is

*A) $\cdot\dot{B}\cdot$

B) $\cdot B\cdot$

C) $\cdot B$

D) $\cdot\dot{B}:$

43) The Lewis dot diagram for sodium with electron configuration $1s^2\ 2s^2\ 2p^6\ 3s^1$ is

*A) $Na\cdot$

B) $Na:$

C) $\cdot\dot{N}a\cdot$

D) $:Na\cdot$

13A.9.

44) The electron configuration for $^{32}_{16}S$ is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$. The Lewis dot diagram for sulfur is

A) ·S·

B) :S:

C) :S:

*D) :S:

45) Given that the Bohr relationship for determination of the energy levels in a hydrogen atom is $E_n = -\frac{1680}{n^2}$ kJ/mol and $h = 4.00 \times 10^{-13}$ kJ·s/mol the frequency (f) of light emitted in an n = 4 to n = 1 transition is (recall E = hf)

A) $2.60 \times 10^{12}\ s^{-1}$

B) $-2.60 \times 10^{12}\ s^{-1}$

*C) $3.94 \times 10^{15}\ s^{-1}$

D) $3.94 \times 10^{-15}\ s^{-1}$

46) The electron configuration of beryllium is $1s^2\ 2s^2$. The number of electrons pictured in a Lewis dot diagram of beryllium is

A) 1

*B) 2

C) 3

D) 4

13A.9.

47) An element is in period 3 and family I of the periodic table. The atoms of this element have

*A) 1 valence and 10 core electrons
B) 3 valence and 6 core electrons
C) 3 valence and 8 core electrons
D) 5 valence and 10 core electrons

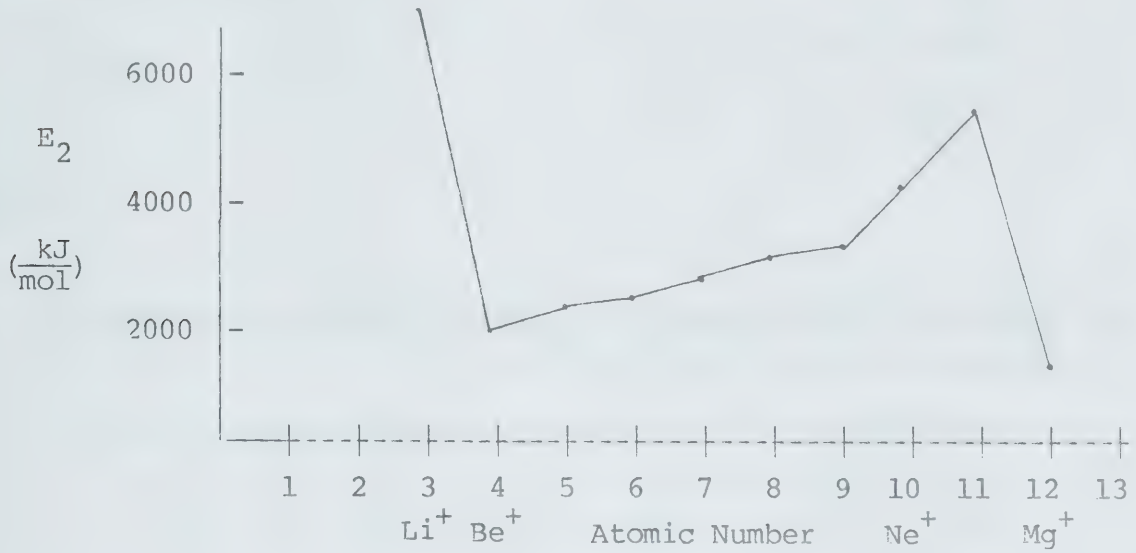
48) What is the name for the energy term represented in the following equation?

$$X^{o} + e^{-} \rightarrow X^{-} + \text{energy}$$

A) entropy
*B) electron affinity
C) ionization energy
D) penetration effect

13A.9.

49) Consider the following plot of the second ionization energy, E_2, for the gaseous ions of the elements from lithium through magnesium.



The general trend in E_2 from Be^+ to Ne^+, is largely due to increasing

*A) nuclear charge
B) density
C) nuclear mass
D) chemical reactivity

50) Which of the following ground state electron configurations will most likely represent an alkali metal?

A) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$

*B) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$

C) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2$

D) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^2$

13A.9.

51) Which one of the following quantum number descriptions specifies a spherical orbital?

A) $n = 2$

B) $m = 0$

C) $m_s = + \frac{1}{2}$

*D) $\ell = 0$

52) The Lewis dot diagram for oxygen with the electron configuration $1s^2 \ 2s^2 \ 2p^4$ is

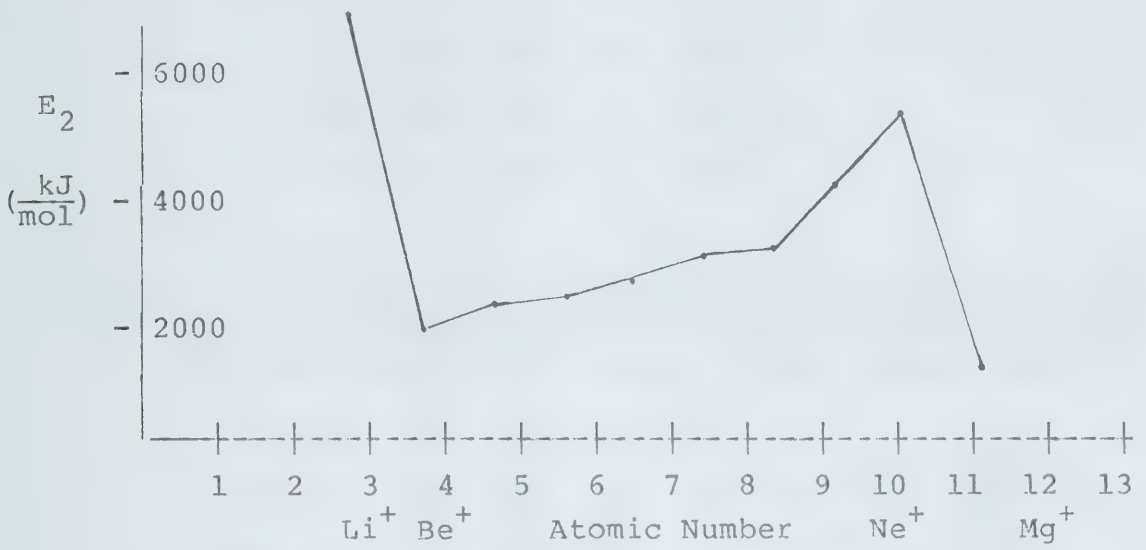
A) :O:

B) ·O·

C) ·O·

*D) :Ö·

53) Consider the following plot of the second ionization energy, E_2, for the gaseous ions of the elements from lithium through magnesium.



A **major reason** why the second ionization energy, E_2, for Li^+ differs as it does from E_2 for Be^+ is that

A) Li^+ has one less neutron in its nucleus than has Be^+.

B) Be^+ has a lower nuclear charge than Li^+.

*C) Be^+ has one 2s electron remaining and Li^+ has no 2s electrons.

D) Be^+ has one more neutron in its nucleus than Li^+.

54) When an electron is in the lowest energy level of an isolated hydrogen atom the quantum number which has a value of 1 is

*A) n

B) ℓ

C) m

D) m_s

13A.9.

55) The maximum number of electrons that can occupy a "d" sub-shell is

A) 6
B) 2
C) 8
*D) 10

56) The principal quantum number of an energy level is n; the maximum total number of orbitals possible (for any value of n) is

A) $(n - 1)^2$
B) $2n^2$
*C) n^2
D) $(n + 1)^2$

57) What is the most common charge on the ions of the Group VI elements?

A) 6-
*B) 2-
C) 6+
D) 4+

58) Given the de Broglie wave equation ($\lambda = \frac{h}{mc}$) which of the following particles will have the shortest wavelength?

A) electron
B) proton
C) neutron
*D) alpha particle

13A.9.

59) The electron configuration that represents a reactive non-metallic element is

A) $1s^2\ 2s^2\ 2p^6\ 3s^1$

B) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$

*C) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

D) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

60) Which one of the following is the correct electron configuration for the element whose atomic number is 32?

A) $1s^2\ 2s^4\ 2p^6\ 3s^6\ 3p^9\ 3d^5$

B) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^2\ 4p^6\ 5s^1$

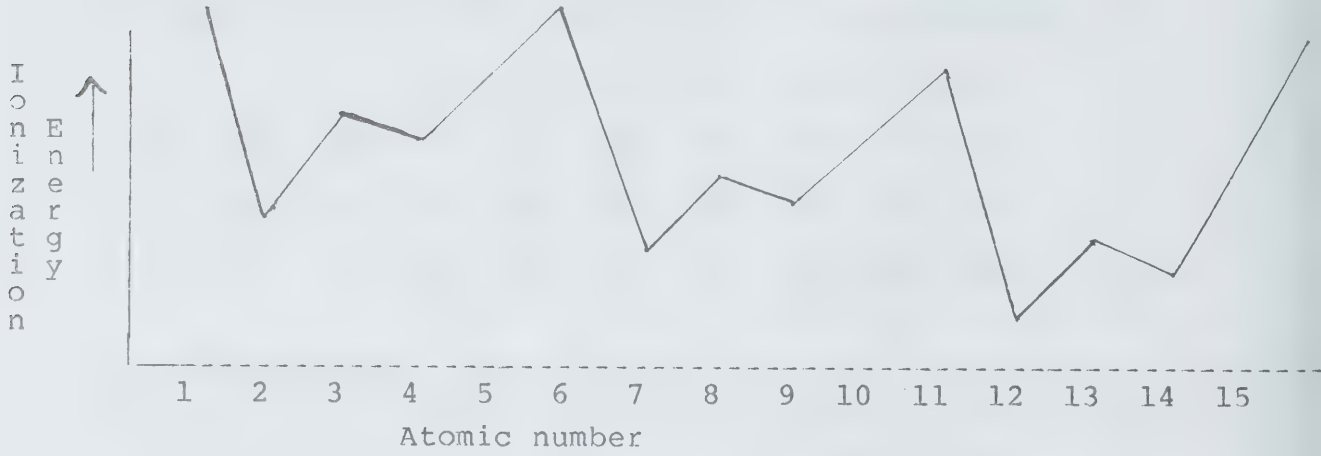
*C) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^2$

D) $1s^2\ 2s^2\ 2p^8\ 3s^2\ 3p^8\ 3d^8\ 4p^2$

13A.9.

61) Given below is a list of fifteen fictitious elements and an experimentally produced graph of their first ionization energies.

Atomic Number	Symbol	Name of Element	Atomic Number	Symbol	Name of Element
1	A	aquarius	9	I	inklium
2	B	butyrium	10	J	jothrium
3	C	clutter	11	K	kringlium
4	D	doople	12	L	luciumm
5	E	effressium	13	M	mentium
6	F	floogium	14	N	noxium
7	G	geranium	15	O	oogle
8	H	hoople			



The formula for the most likely compound between geranium and jothrium is

*A) GJ

B) G_2J

C) G_3J_2

D) GJ_2

62) When the quantum number ℓ = 0, m may take on the values

A) 0, 1, 2, 3 ...

B) 1 only

C) 1, 2, 3, 4 ...

*D) 0 only

13A.9.

63) The Noble gases are characterized by completed outer shells of electrons. The outer shells of the Noble gases - with the exception of helium - consist of

A) s and d electrons
*B) s and p electrons
C) p and d electrons
D) s, p and d electrons

64) The electronic configuration of the fluoride ion, $^{19}_{9}F^{-}$, may be represented by

A) $1s^2\ 2s^2\ 2p^5$
B) $1s^2\ 2s^2\ 2p^6\ 3s^1$
C) $1s^2\ 2s^2\ 2p^6\ 3s^2$
*D) $1s^2\ 2s^2\ 2p^6$

65) The total number of orbitals possible in the third energy level (n=3) of any atom is

A) 8
*B) 9
C) 3
D) 18

66) A metallic element, M, has the electronic configuration 2, 8, 2. The most stable ion of element M would be best represented by

A) M^{3+}
B) M^{2-}
*C) M^{2+}
D) M^{1+}

l3A.9.

67) Which of the following is the element in the third period of the Periodic Table having only one "p" electron in its valence shell?

A) Mg
B) Ge
*C) Al
D) C

68) An atomic orbital is best described as

A) a volume in space
B) an electron path
*C) a probability function
D) a set of concentric circles

69) As a general statement, for the electrons in the n^{th} level of any atom (for the 4th energy level, n = 4) the number of orbitals is equal to

A) n-1
B) n
*C) n^2
D) $2n^2$

70) The atomic orbitals in order of **increasing** energy are

*A) s < p < d
B) d < s < p
C) s < d < p
D) p < s < d

13A.9.

71) The orbitals of "p" electrons

A) are spherical about the nucleus
*B) lie along each of three coordinate axes
C) have any number of electrons in them
D) are found only in the first energy level (n = 1)

72) The electron configuration of an excited neon atom may be represented by

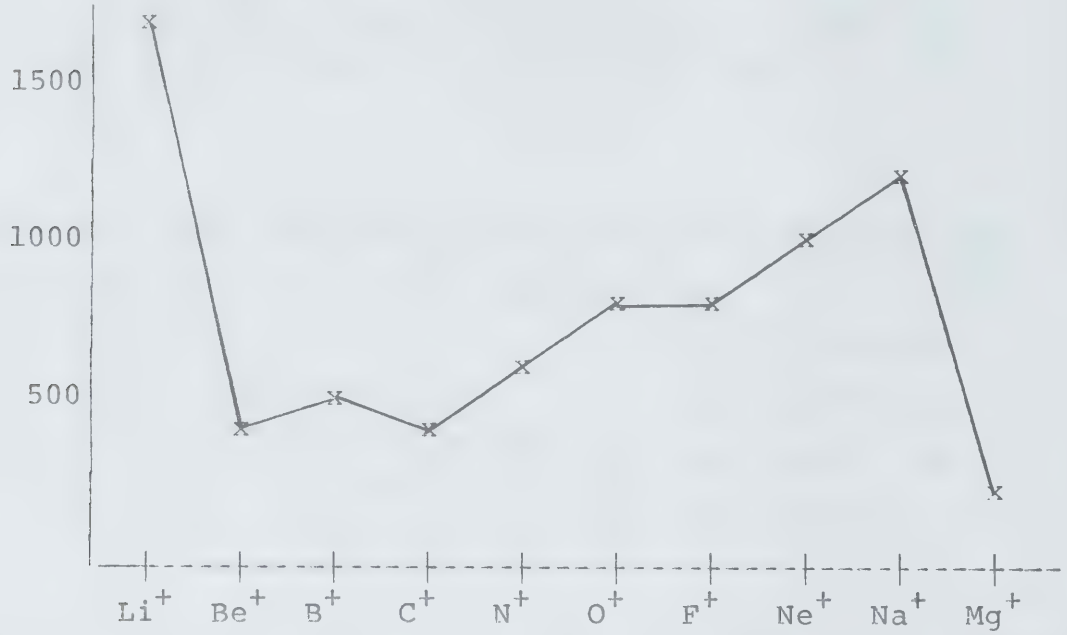
A) $1s^2\ 2s^2\ 2p^6$
B) $1s^2\ 2s^2\ 2p^5$
*C) $1s^2\ 2s^2\ 2p^5\ 3s^1$
D) $1s^2\ 2s^2\ 2p^6\ 3s^1$

73) The photoemission of electrons is evidence for the

*A) particle theory of light
B) wave theory of light
C) wave theory of electrons
D) line spectra of atoms

13A.9.

74) Consider the figure below which shows a plot of the second ionization energy E_2 for gaseous ions of the elements from lithium through magnesium



Of the following statements, which best corresponds with the difference between E_2 for Li^+ and E_2 for Be^+?

A) Lithium is softer than is beryllium.
B) Lithium has a lower melting point than has beryllium.
C) Lithium is a better conductor of electricity than is beryllium.
*D) Lithium forms LiF whereas beryllium forms BeF_2

75) In the third energy level of a many electron atom, the type(s) of orbitals possible are

A) s
B) s, p
*C) s, p, d
D) s, p, d, f

13A.9.

76) Which of the following species is isoelectronic with argon?

A) Ne
B) Cl^+
C) P^{3+}
*D) S^{2-}

77) The electron configuration $1s^2\ 2s^2\ 2p^6\ 3s^2$ is representative of a neutral atom with

*A) atomic number 12
B) 8 valence electrons
C) atomic number 8
D) relative atomic mass of 12

78) An element has the following ionization energies (kJ/mol)

E_1	E_2	E_3	E_4
900	1750	14785	20947

The element is most likely

A) Bi
*B) Be
C) Bk
D) Br

13A.9.

79) The energies of the orbitals of the hydrogen atom are represented by:

$$E_n = (-1317/n^2) \text{ kJ/mol.}$$

To excite an electron from a 1s to a 3s orbital, the energy required is

A) 146.3 kJ/mol
B) 342.8 kJ/mol
*C) 1171 kJ/mol
D) 1317 kJ/mol

80) What is the general trend in the ionization energy down a vertical group (Li, Na, K, Rb, Cs) in the Periodic Table?

A) increases
*B) decreases
C) remains unchanged
D) behaves unpredictably

81) The energy difference between two electron orbitals is 5.22 kJ/mol.

$$(h = 6.49 \times 10^{-37} \text{ kJ·s})$$

The frequency of light emitted when the electron drops from the higher to the lower state is

A) 1.24×10^{37} Hz
B) 7.47×10^{14} Hz
*C) 1.34×10^{13} Hz
D) 8.00×10^{36} Hz

13A.9.

82) Which of the following statements about energy levels of atoms are correct?

I An atom is in its ground state when it has its least possible amount of energy.

II The atom in the ground state may absorb only those quanta that will enable its electron(s) to move to a higher energy level.

III The atoms can absorb any amounts of energy whatever.

IV The energy levels of electrons in atoms can be described by quantum numbers.

A) I, II and III

*B) I, II and IV

C) II and IV

D) I and IV

83) The lowest second ionization energy would be expected for the element with the electron configurations

*A) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$

B) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$

C) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

D) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$

84) What is the highest occupied electronic subshell in a chlorine atom in its ground state?

A) 2p

B) 3s

*C) 3p

D) 4p

85) The ionization energy of oxygen (1318 kJ/mol) is lower than the ionization energy of nitrogen (1407 kJ/mol). This is mainly due to

A) an increase in the nuclear charge as atomic number increases from 7 to 8

B) the increased distance between the valence electrons and the nucleus in oxygen

*C) the extra repulsion from the 8th electron entering the partially filled $2p_x$ orbital of oxygen

D) the smaller size of the oxygen atom

86) Wave particle duality theory is associated mainly with which of the following men?

A) Schrödinger

B) Rutherford

C) Bohr

*D) De Broglie

87) Consider the following electron configurations:

I $1s^2\ 2s^2\ 3s^2$

II $1s^2\ 2s^2\ 2p^6$

III $1s^2\ 2s^2\ 2p^6\ 3s^1$

IV $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^8$

V $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6$

The configurations that represent elements of the same group in the periodic table are

A) IV and V

B) I and IV

*C) II and V

D) II and III

l3A.9.

88) The element whose atoms are is best represented by the electron configuration

$$1s^2\ 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^1$$

is

A) carbon
B) oxygen
*C) nitrogen
D) boron

89) Within a family in the periodic table, why is the metal with the lowest ionization energy usually the one with the highest atomic number?

A) Larger atoms are less likely to be spherical.
B) Larger atoms are less likely to pack close together.
C) The heaviest particles have the greatest kinetic energies.
*D) The distance between the nucleus and the electrons is greatest for the heaviest atoms.

90) Which of the following species could have the electronic configuration $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^0\ 3d^8$?

A) Fe^+
B) Ni
*C) Ni^{2+}
D) Cu^{2+}

91) An orbital is considered by modern scientists to be
A) a spherical or dumbbell-shaped pathway, outside the nucleus, followed by an electron
B) a region in space, having a precise shape, which is completely filled by the extremely dense electron
*C) a region in space in which the probability of finding an electron is high
D) the same as an orbit, but with no trajectory

13A.9.

92) Which of the following represents the electron configuration of an excited atom?

A) $1s^2\ 2s^2\ 2p^6$

B) $1s^2\ 2s^1$

*C) $1s^2\ 2s^2\ 2p^2\ 3s^1$

D) $1s^2\ 2s^2\ 2p^6\ 3s^1$

93) The electron configuration for a sulphide ion S^{2-} is

A) $1s^2\ 2s^2\ 2p^6$

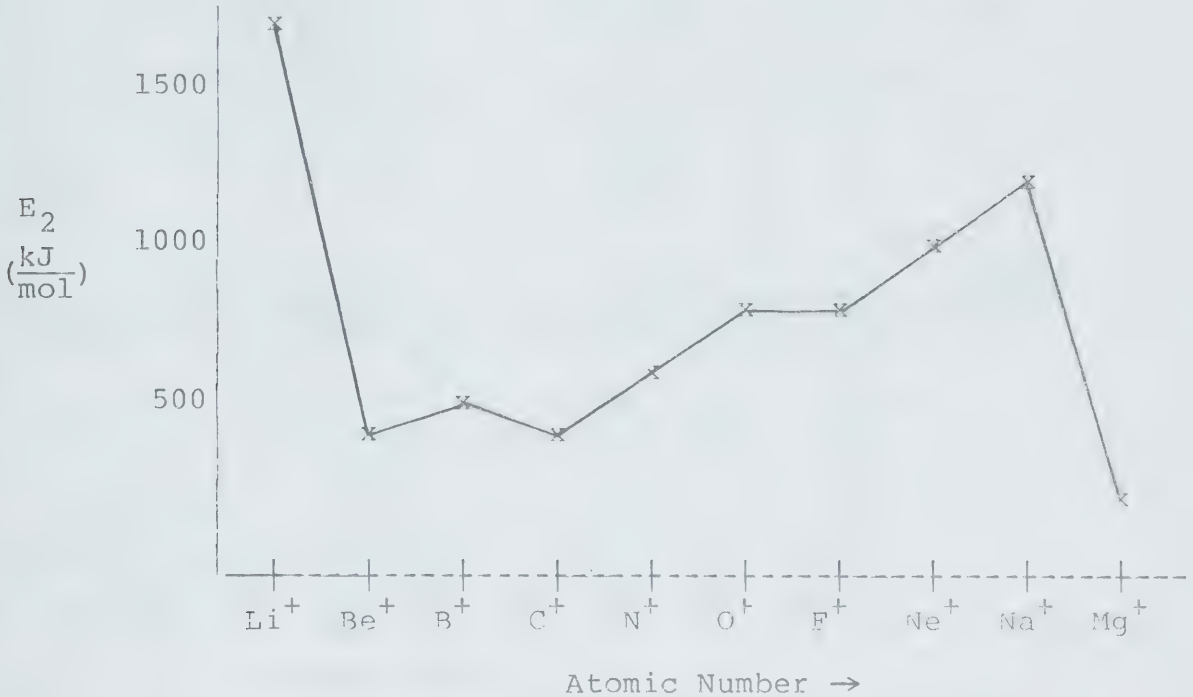
B) $1s^2\ 2s^2\ 2p^6\ 3s^0\ 3p^6$

*C) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

D) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$

13A.9.

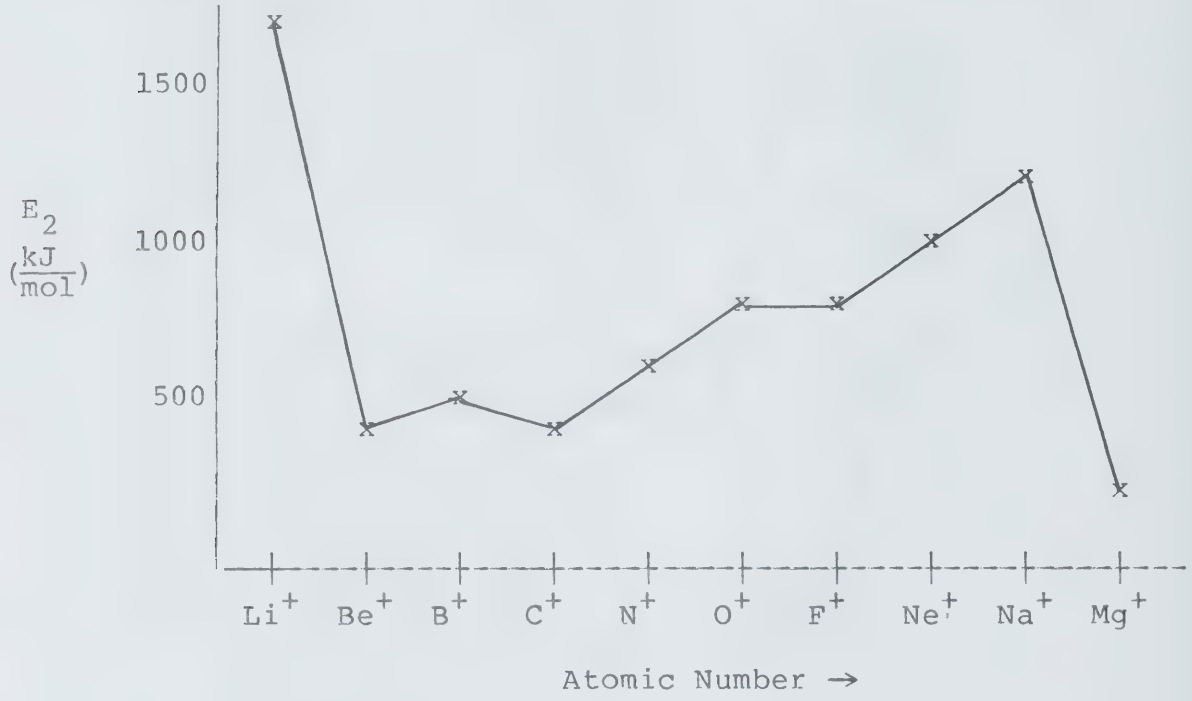
94) The figure below shows a plot of the second ionization energy, E_2, for gaseous ions of elements from lithium through magnesium.



The general trend in E_2 from Be^+ to Ne^+ is largely due to

A) nuclear mass
*B) nuclear charge
C) metallic character
D) chemical reactivity

95) The figure below shows a plot of the second ionization energy, E_2, for gaseous ions of elements from lithium through magnesium.



One reason why the second ionization energy E_2 for Li^+ differs from E_2 for Be^+ is that

A) Li^+ has one neutron less in its nucleus than Be^+

B) Be^+ has a lower nuclear charge than Li^+ has

*C) Be^+ has one 2s electron remaining and Li^+ has no 2s electrons

D) Lithium has more naturally occurring isotopes than beryllium has

13A.10.

1) The molecular shape resulting from an sp^3 combination of bonding orbitals in which all four orbitals form chemical bonds is

A) pyramidal
*B) tetrahedral
C) planar triangular
D) bent

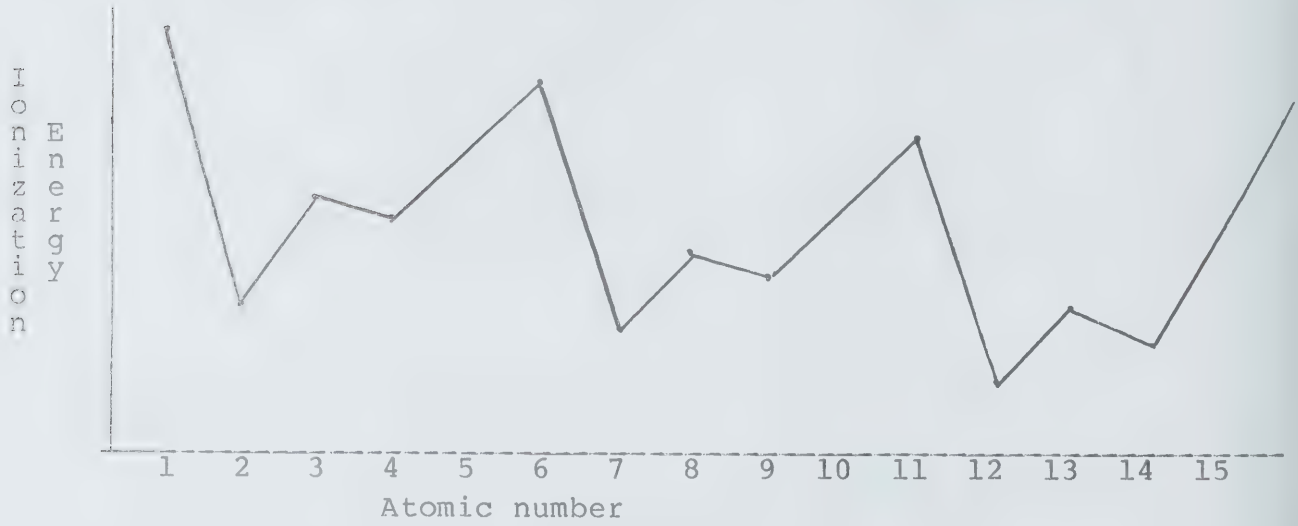
2) The greatest degree of ionic bonding would be expected to occur between atoms with electron configurations

A) $1s^2\ 2s^2\ 2p^1$ and $1s^2\ 2s^2\ 2p^5$
B) $1s^2\ 2s^2\ 2p^1$ and $1s^2\ 2s^2\ 2p^6\ 3s^2$
*C) $1s^2\ 2s^2\ 2p^5$ and $1s^2\ 2s^2\ 2p^6\ 3s^1$
D) $1s^2\ 2s^2\ 2p^1$ and $1s^2\ 2s^2\ 2p^6\ 3s^1$

13A.10.

3) Listed below are fifteen fictitious elements and an experimentally produced graph of their first ionization energies.

Atomic Number	Symbol	Name of Element	Atomic Number	Symbol	Name of Element
1	A	aquarius	9	I	inklium
2	B	butyrium	10	J	jothrium
3	C	clutter	11	K	kringlium
4	D	doople	12	L	lucium
5	E	effressium	13	M	mentium
6	F	floogium	14	N	noxium
7	G	geranium	15	O	oogle
8	H	hoople			



The type of compound formed between geranium and jothrium will most likely be a(n)

*A) ionic solid
B) network solid
C) covalent molecule
D) diatomic gas

4) Which of the compounds listed below is the BEST example of a network solid?

A) NaCl(s), table salt

B) $C_{25}H_{52}$, paraffin

C) $S_8(s)$, sulphur

*D) $SiO_2(s)$, quartz

5) From the energy data listed below, which of the crystals would most likely be a molecular crystal?

Crystal	Energy Required to Separate the Particles in the Crystal
A)	6.55 kJ/mol
B)	415.8 kJ/mol
C)	714.0 kJ/mol
*D)	2511.6 kJ/mol

6) The alkali metals are softer than most other metals because

A) their atomic radii are smaller

B) they are more electropositive

C) they are closer to the noble gas configurations

*D) they have fewer valence electrons

7) The type of bonding in CO_3^{2-} (carbonate ion) is best described by which one of the following terms?

*A) covalent

B) Van der Waal's

C) ionic

D) hydrogen

13A.10.

8) Which of the following is the most fundamental definition of a chemical bond?

*A) Electrons simultaneously attracted by more than one nucleus.

B) Positive ions attracting negative ions.

C) Transfer of electrons from one atom to another.

D) Attraction between valence electrons from two atoms.

9) Which of the following is the correct Lewis dot diagram for PH_3?

A) H·P·H
   ·
   H

*B) H:P:H
   ..
   H

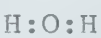
C) H:P:H
   ..
   H

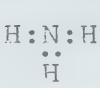
D) H:P:H
   ..
   H

13A.10.

10) Which of the following Lewis diagrams is a correct representation?

*A)

B) 

C) 

D)

11) The valence of element X is 2 and the valence of element Y is 3. The formula of the compound produced when they unite is

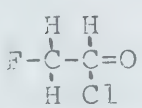
A) 2X3Y

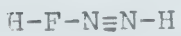
B) X_2Y_3

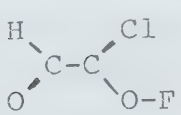
*C) X_3Y_2

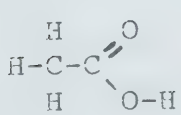
D) X_3Y

12) Which of the following structural formulae is correct?

A) 

B) 

C) 

*D) 

13A.10.

13) The bonding capacity of boron is

A) 1
B) 2
*C) 3
D) 5

14) Given that element X will form diatomic molecules with each of the elements P,Q,S,T and that the ionization energies of the elements are:

X = 588 kJ/mol
P = 1008 kJ/mol
Q = 840 kJ/mol
S = 1764 kJ/mol
T = 1344 kJ/mol

Which of the following lists the compounds in increasing order of ionic nature?

A) X-P, X-Q, X-S, X-T
*B) X-Q, X-P, X-T, X-S
C) X-Q, X-T, X-P, X-S
D) X-S, X-T, X-P, X-Q

15) Nitrogen is a gas at temperatures above 126 K. Very little energy is required to change nitrogen from solid to liquid to gas. However, temperatures of thousands of degrees Kelvin do not cause appreciable decomposition of the N_2 molecules. What forces or bonds are primarily responsible for holding the two nitrogen atoms together in a molecule of nitrogen gas, N_2?

A) ionic bonds
B) hydrogen bonds
*C) covalent bonds
D) van der Waals forces

13A.10.

16) Which one of the following pairs of compounds consists of only polar molecules?

A) CCl_4, CH_3Cl
B) NH_3, CCl_4
*C) NH_3, H_2S
D) H_2O, BF_3

17) Which of the following molecules is the least polar?

A) H-F
*B) O=C=O
C)
D)

18) The shape of the carbonate ion, CO_3^{2-}, is

A) linear
*B) trigonal planar
C) pyramidal
D) tetrahedral

19) Which pair of elements will form a compound having bonds with the most ionic character?

A) lithium and boron
B) carbon and fluorine
*C) potassium and chlorine
D) chlorine and bromine

13A.10.

20) The sharing of electrons in bond formation always involves

A) formation of positive and negative ions
B) formation of polar molecules
C) shared electrons being attracted more by one atom than another
*D) lower energy content for the bonded atoms than for the unbonded atoms

21) Which of the following bonds is the most polar?

A) C-O
B) C-N
*C) B-O
D) B-N

22) Which compound exhibits the greatest covalent character?

*A) NO
B) MgO
C) NaCl
D) RbS

23) Which of the orbitals in the group I elements of the Periodic Table (alkali metals family) contains the "bonding electron"?

*A) s
B) p
C) sp
D) d

24) Which one of the molecules listed has a triple covalent bond?

A) NH_3
B) O_2
C) F_2
*D) N_2

25) The shape of a BF_3 molecule is

A) tetrahedral
*B) triangular planar
C) octahedral
D) square pyramidal

26) Which of the following would be expected to produce the strongest electrolyte in aqueous solution?

A) CCl_4
B) CS_2
C) I_2
*D) CaF_2

27) Which of the following substances has a molecular dipole moment of zero?

A) NH_3
B) CH_2Cl_2
C) $CHCl_3$
*D) CCl_4

13A.10.

28) Which of the following compounds consists of tetrahedral molecules?

A) CsCl
B) CO_2
*C) CCl_4
D) H_2O

29) The F-B-F angle in a BF_3 molecule is

A) 90°
B) 102°
C) 109.5°
*D) 120°

30) The shape of an NF_3 molecule is

A) tetrahedral
B) trigonal bipyramidal
C) octahedral
*D) pyramidal

31) Which of the following molecules would have a net dipole of zero?

I H_2O
II $CHCl_3$
III CCl_4
IV CO_2
V NH_3
VI BF_3

A) II, III
B) III, IV
C) III, IV, V
*D) III, IV, VI

32) What type of bonding exists between phosphorus and hydrogen in the compound PH_3?

A) ionic bonding
*B) covalent bonding
C) hydrogen bonding
D) van der Waals forces

33) Which of the following compounds has the highest degree of ionic bonding?

*A) CsCl
B) CO_2
C) CCl_4
D) H_2O

34) Which of the following compounds has double bonds within its molecular structure?

A) CsCl
*B) CO_2
C) CCl_4
D) H_2O

35) Which of the formulae represents compounds whose molecules are tetrahedral?

A) BH_3
B) C_2H_2
*C) CH_4
D) H_2O

13A.10.

36) In which of the following compounds is sp^2 hybridization used to explain the bonding?

*A) BH_3
B) C_2H_2
C) CH_4
D) H_2O

37) Atoms **within a molecule** are held together by which one of the following forces?

A) gravitational
*B) electric
C) magnetic
D) van der Waals

38) The correct Lewis dot diagram for the water molecule is

A) H:$\ddot{O}$:H

B) H:O:H

*C) H:$\ddot{O}$:H
$\ddot{}$

D) H:$\dot{O}$:H
$\dot{}$

39) Of the elements listed, the element which forms the **most** covalent bond with fluorine is

A) Li
*B) Be
C) K
D) Ca

40) A coordinate covalent bond is one in which

*A) one atom shares two of its electrons with another atom

B) two atoms each contribute one electron to a common bond

C) two atoms contribute two electrons each to a common bond

D) two atoms are bonded by one electron

41) What type of bonding is found between adjacent molecules of oxygen?

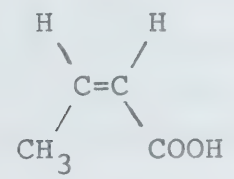
A) ionic

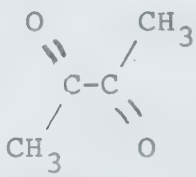
B) covalent

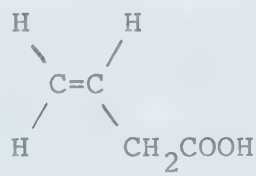
*C) Van der Waals

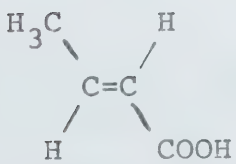
D) coordinate covalent

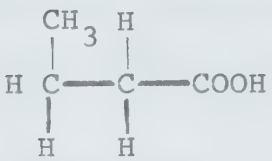
42) Consider five possible structures of the molecule $C_4H_6O_2$.

I) 

II) 

III) 

IV) 

V) 

Cis-trans isomerism is illustrated by which pair of structures?

A) I and II

*B) I and IV

C) III and IV

D) III and V

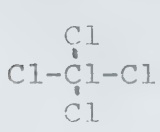
13A.10.

43) An example of a polar molecule is represented by the formula

*A) H-Cl

B) O=C=O

C) N≡N

D) 

44) The shape of the molecule BCl_3 (boron(III) chloride) is

A) linear
B) octahedral
*C) trigonal planar
D) trigonal pyramidal

45) A molecule of CH_4 (methane) is

A) linear and nonpolar
B) V-shaped and polar
*C) tetrahedral and nonpolar
D) pyramidal and polar

46) A molecule of NH_3 (ammonia) is

A) linear and polar
B) V-shaped and polar
C) tetrahedral and nonpolar
*D) pyramidal and polar

13A.10.

47) The shape of the CO_2 (carbon dioxide) molecule is

A) pyramidal
B) trigonal bipyramidal
C) octagonal
*D) linear

48) The shape of CH_2Cl_2 (dichloromethane) is most likely

*A) tetrahedral
B) linear
C) planar
D) pyramidal

49) When a carbon atom is bonded to four hydrogen atoms in CH_4 (methane), the shape of the molecule is described as

A) square
*B) tetrahedral
C) rectangular
D) octahedral

50) The abnormally high boiling point of water is due primarily to the presence of

A) covalent bonding
B) dipole-dipole attractions
C) Van der Waals attractions
*D) hydrogen bonding

13A.10.

51) The substance listed below that has polar molecules, predominantly covalent bonding between atoms within the molecules and a significant degree of hydrogen bonding between molecules in the liquid state is

*A) HF (hydrogen fluoride)

B) Na_2O (sodium oxide)

C) PH_3 (phosphine)

D) SiH_4 (silane)

52) Which one of the following atom pairs would form the most polar covalent bond?

A) F and F

B) N and O

*C) H and F

D) H and Br

53) The bond in HCl (hydrogen chloride) is said to be polar covalent because

A) hydrogen chloride molecules are relatively weakly held together

*B) the shared electron pair is closer to the one atom than to the other

C) chlorine has a much higher atomic number than hydrogen

D) chemical bonds consist of electrical forces between atoms

54) Given the compounds:

I	CH_4	(methane)
II	H_2S	(hydrogen sulphide)
III	CH_3F	(fluoromethane)
IV	BCl_3	(boron trichloride)
V	PCl_5	(phosphorus pentachloride)

Which of the following groups contain only polar compounds?

A) II and V
B) III and V
C) I, IV and V
*D) II and III

55) The H-N-H bond angle in ammonia (NH_3) is less than the H-C-H angle in methane (CH_4) due to

A) repulsions between hydrogen atoms in ammonia
B) attractions between hydrogen atoms in methane
*C) a non-bonded pair of electrons in ammonia
D) the tetrahedral shape of the molecules in ammonia and methane

56) The bonding orbitals on the cental atom in a BF_3 (boron trifluoride) molecule are best described as

A) s orbitals
B) sp orbitals
*C) sp^2 orbitals
D) sp^3 orbitals

57) The shape of a water molecule is best described as

A) planar
*B) bent
C) pyramidal
D) tetrahedral

58) In neutral molecules the bonding capacity of an oxygen atom is "2" because

A) there are 2 valence electrons in the 2s orbital
B) there are 2 types of orbitals with principal quantum number 2
C) oxygen is in the 2nd row of the periodic table
*D) oxygen has 6 valence electrons

59) Which of the following compounds exhibits the greatest covalent character?

*A) NO (nitrogen monoxide)
B) MgO (magnesium oxide)
C) NaCl (sodium choride)
D) RbS (rubidium sulphide)

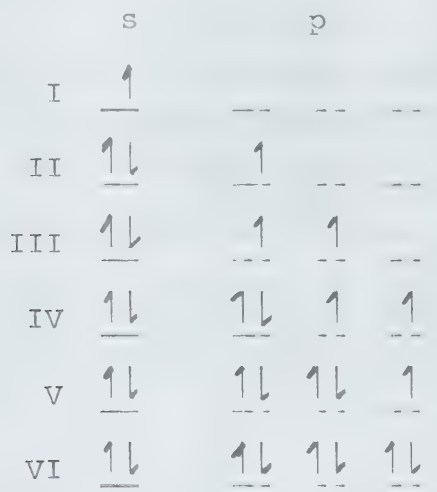
60) Which of the following statements is true?

*A) CCl_4 (carbon tetrachloride), at room temperature, is a liquid with a negligible dipole moment.
B) NH_3 (ammonia) is an example of a planar molecule.
C) SF_6 (sulphur hexafluoride) has a tetrahedral structure.
D) BCl_3 (boron trichloride) is polar.

61) The bond of H^+ to NH_3 in NH_4^+ is

A) ionic
*B) coordinate covalent
C) van der Waals
D) metallic

62) Consider the following diagram representing electron configurations for the outer energy levels of atoms of different elements:



Which of the following combinations of the above elements would be most likely to form ionic bonds?

A) I and II
*B) I and V
C) III and IV
D) IV and V

63) Which of the following has no net dipole?

A) KBr (potassium bromide)
B) $CHCl_3$ (trichloromethane)
C) HF (hydrogen fluoride)
*D) C_2H_2 (acetylene)

64) Which of the following statements about chemical bonding is true?

A) A chemical bond can result only when the sharing of electrons between two or more atoms is equal.
B) The strongest chemical bonds result when antibonding orbitals are used.
C) Atoms combine to form bonds with the absorption of energy.
*D) Chemical bonds can form when empty or half filled orbitals permit electrons to position themselves between nuclei.

65) The first ionization energies for O, Cl, Br and P are 1.32×10^3 kJ/mol, 1.26×10^3 kJ/mol, 1.13×10^3 kJ/mol and 1.02×10^3 kJ/mol respectively. Which pair of elements would form the most polar bond?

A) O and Cl
B) Cl and Br
*C) O and P
D) P and Br

66) Given that the following elements are arranged in order of their increasing ability to attract the electrons in a "covalent" bond:

H C Cl F

The compound which is most polar is

A) $CHCl_3$
B) CCl_2F_2
*C) CHF_3
D) CCl_3F

13A.10.

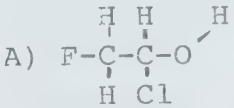
67) Element Y forms compounds A-Y, B-Y, C-Y and D-Y. The elements Y, A, B, C and D have the following first ionization energies (E_1):

Y	840 kJ/mol
A	672 kJ/mol
B	1050 kJ/mol
C	798 kJ/mol
D	1302 kJ/mol

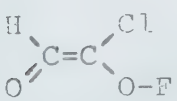
The compounds in order of increasing ionic character of the bond are

*A) C-Y A-Y B-Y D-Y
B) C-Y A-Y D-Y B-Y
C) B-Y D-Y A-Y C-Y
D) D-Y B-Y A-Y C-Y

68) Which of the following structural formulas is **incorrect** according to the bonding capacities of the atoms concerned?

A) 

B) F-N=N-H

*C) 

D) N≡N

69) The shape of the SF_6 molecule is likely to be

*A) octahedral
B) linear
C) tetrahedral
D) square planar

70) Which of the following molecules is most polar?

*A) HCl (hydrogen chloride)
B) CO_2 (carbon dioxide)
C) I_2 (iodine)
D) SF_6 (sulphur hexafluoride)

71) The bonding orbitals on the central atom in a CF_4 (carbon tetrafluoride) molecule is are described as

A) p orbitals
B) sp orbitals
*C) sp^3 orbitals
D) sp^2 orbitals

72) The shape of an NH_3 (ammonia) molecule is

A) linear
B) planar
*C) pyramidal
D) tetrahedral

73) The valence electron configuration in atoms of element X is $2s^2\ 2p^4$ and in those of element Z is $4s^2\ 4p^5$. The probable empirical formula for the compound formed from these elements is

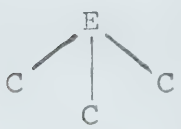
A) XZ
B) X_2Z
*C) XZ_2
D) X_2Z_3

13A.10.

74) Consider the following elements and their first ionization energies

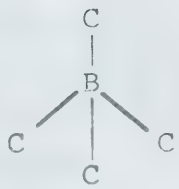
A 540 kJ/mol
B 630 kJ/mol
C 840 kJ/mol
D 460 kJ/mol
E 710 kJ/mol

The following 3 compounds are formed



I (Pyramidal)

II (Pyramidal)



III (Tetrahedral)

Which of the following statements concerning the three molecules is true?

A) Electrons are attracted more strongly to B than to C.
*B) I and II are polar while three is non-polar.
C) Molecules I and II have the same structure (pyramidal) and therefore have the same degree of polarity.
D) The bond angles in molecules I and II would be identical.

75) The simplest formula for a compound between elements A ($1s^2\ 2s^2\ 2p^6\ 3s^2$) and B ($1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$) is

A) AB
*B) AB_2
C) A_2B
D) AB_3

13A.11.

1) Nitrogen is a gas at temperatures above 126 K. Which forces or bonds between nitrogen molecules are primarily responsible for maintaining the nitrogen in the liquid state?

A) ionic bonds
B) hydrogen bonds
C) covalent bonds
*D) van der Waals forces

2) Which one of the following is classed as an ionic solid?

*A) NaBr(s)
B) CO_2(s)
C) SO_2(s)
D) HCl(s)

3) The substance S_8 (sulfur), with a melting point of 392 K is an example of

A) an ionic solid
B) a metallic solid
*C) a molecular solid
D) an amorphous solid

4) The major bonding or force that exists in liquid helium is

*A) a van der Waals force
B) polar covalent bonding
C) metallic bonding
D) nonpolar covalent bonding

13A.11.

5) Water has a higher boiling point than hydrogen sulfide because water has

A) smaller molecules
B) a smaller molecular mass
C) weaker van der Waals forces
*D) stronger intermolecular attractions

6) Van der Waals forces become strongest in gases at

A) high temperatures and high pressures
B) high temperatures and low pressures
*C) low temperatures and high pressures
D) low temperatures and low pressures

7) Van der Waals forces will be the strongest between molecules that have a

*A) large number of electrons and are a small distance apart
B) large number of electrons and are a great distance apart
C) small number of electrons and are a great distance apart
D) small number of electrons and are a small distance apart

8) Fluorine and chlorine are gases, bromine is a liquid, and iodine is a solid at room temperature. Therefore, van der Waals forces must be greatest between the molecules of

*A) iodine
B) bromine
C) chlorine
D) fluorine

13A.11.

9) If Ca^{2+} (calcium ion) combines with CO_3^{2-} (carbonate ion), the result will be

A) a metallic crystal
B) a molecular crystal
C) a covalent network solid
*D) an ionic crystal

10) Solid iodine sublimes readily upon heating. This is an indication of what type of bonds in the iodine solid?

A) ionic bonds
*B) van der Waals forces
C) coordinate covalent bonds
D) metallic bonds

11) The type of solid that contains particles held together primarily by van der Waals forces is

A) ionic
B) metallic
C) network
*D) molecular

12) Many salts dissociate in water because water

*A) consists of polar molecules
B) has a linear structure
C) contains ionic bonds
D) does not ionize

13A.11.

13) Potassium chloride dissolved in pure water is an example of

*A) an ionic compound dissolved in a polar solvent
B) an ionic compound dissolved in a nonpolar solvent
C) a covalent compound dissolved in a polar solvent
D) a covalent compound dissolved in a nonpolar solvent

14) Which one of the following statements is true?

A) An element with an electronic configuration $1s^2 2s^2 2p^5$ has five valence electrons.
B) F_2 is a polar molecule.
*C) HF is a polar molecule.
D) Fluorine is the least electronegative element.

15) Metals can be rolled into sheets and stamped into various forms whereas diamond is very hard and brittle. An explanation of these properties is that

A) metals have covalent bonds, while diamond has ionic bonds in which electrons are transferred between neighbouring carbon atoms
*B) electrons surrounding a given metal atom are free to move throughout the metal, while in a diamond, the bonding electron pairs are held tightly between two carbon atoms in a tetrahedral pattern
C) electrons of metals are held more tightly to the parent atom than are electrons of carbon atoms, hence bonds in metals are always stronger
D) diamond has strong double bonds between carbon atoms, while metal bonds are normal single covalent bonds which bend easily

13A.11.

16) Which of these compounds is the best example of a molecular solid held together primarily by van der Waals forces?

A) KI(s)

B) $Na_2O(s)$

C) $H_2O(s)$

*D) $CO_2(s)$

17) The strength of intermolecular "bonds", due to van der Waals forces, increases in a given family (such as the halogens) as

*A) the number of electrons in the atom increases

B) the molecular size decreases

C) the temperature increases

D) molecules become more nearly spherical

18) In which of the following compounds is hydrogen bonding strongest in the liquid state?

*A) HF

B) CH_4

C) H_2S

D) PH_3

19) When sodium nitrate ($NaNO_3(s)$) dissolves in water, the reaction is endothermic. The endothermicity can be accounted for by assuming

*A) high lattice energy and low hydration energy

B) high lattice energy and high hydration energy

C) low lattice energy and high hydration energy

D) low lattice energy and low hydration energy

20) A pure substance melts at 113°C and does not conduct electricity in either the solid or liquid state. The bonding in this substance is primarily

A) ionic
B) covalent network
C) metallic
*D) covalent molecular

21) What type of bonding holds iodine atoms together in an individual molecule of iodine, $I_2(s)$?

A) ionic bonding
*B) covalent bonding
C) hydrogen bonding
D) van der Waals forces

22) What type of bonding most probably occurs between two elements if one of the elements has a low electronegativity and the other element has a high electronegativity?

*A) ionic bonding
B) covalent bonding
C) hydrogen bonding
D) van der Waals forces

23) What type of substance melts at a temperature greater than 700°C, is extremely hard, not workable and a poor conductor of heat and electricity in both the solid and liquid states?

A) ionic crystal
B) metallic crystal
*C) network crystal
D) molecular crystal

24) Crystals of sodium chloride (NaCl) consist of

A) molecules of NaCl held together by ionic bonds
*B) ions of Na^+ and Cl^- held together by ionic bonds
C) atoms of Na and Cl held together by polar covalent bonds
D) atoms of Na and Cl held together by ionic bonds

25) Network solids

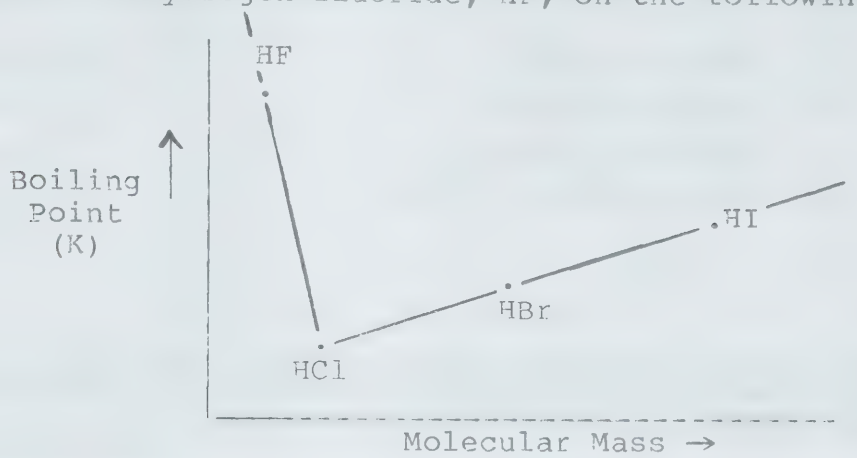
A) are good electrical conductors
*B) have high melting points
C) are very malleable
D) have low boiling points

26) A certain solid consists of ions bound in a crystal lattice. Which property would <u>not</u> be expected to be exhibited by the solid?

A) high melting point
B) hardness and brittleness
*C) good conductivity
D) high solubility in polar solvents

13A.11.

27) Which type of intermolecular bonding best explains the position of hydrogen fluoride, HF, on the following graph?



A) ionic interaction
*B) hydrogen bonding
C) covalent bonding
D) van der Waals forces

28) Carbon and silicon are group IV elements and form the oxides CO_2 and SiO_2 respectively. CO_2 is a gas at room temperature and pressure, while SiO_2 melts at about 1700°C. The great difference in melting points is best accounted for by the fact that

A) CO_2 has non-polar molecules, but SiO_2 has polar molecules
*B) $CO_2(s)$ is a molecular solid, but $SiO_2(s)$ is a covalent or network solid
C) bonds between silicon and oxygen in SiO_2 molecules are many times stronger than bonds between carbon and oxygen in CO_2 molecules
D) silicon-silicon bonds are stronger than carbon-carbon bonds

13A.11.

29) What type of intermolecular bonding holds the many molecules of iodine (I_2) together in a crystal of iodine?

A) ionic bonding
B) covalent bonding
C) hydrogen bonding
*D) van der Waals forces

30) What type of bonding exists between individual water molecules but does not exist between the elements hydrogen and oxygen within a molecule of water?

A) ionic bonding
B) covalent bonding
*C) hydrogen bonding
D) electrovalent bonding

31) What type of substance could melt at 660°C, not conduct electricity in the solid form, but conducts electricity in the liquid state?

*A) ionic crystal
B) metallic crystal
C) covalent network crystal
D) molecular crystal

32) What type of substance could melt at -7.2°C and not conduct electricity in either the solid or liquid state?

A) ionic crystal
B) metallic crystal
C) covalent network crystal
*D) molecular crystal

13A.11.

33) Which of the following compounds exhibits the highest degree of hydrogen bonding between individual molecules?

A) H_2S
B) C_2H_2
C) CH_4
*D) H_2O

34) The abnormally high boiling and melting points of water compared with hydrides of other group VI elements are attributed to

*A) hydrogen bonding
B) van der Waals forces
C) the angular shape of the molecules
D) the high density of water

35) Which of the bonds between the following elements would have the most covalent nature?

A) two metals with high electronegativity difference
B) a metal and non metal with high electronegativity difference
C) a metal and non metal with a low electronegativity difference
*D) two non metals of very low electronegativity difference

36) The substance in the following list that exhibits the greatest degree of hydrogen bonding is

A) CH_4 (methane)
B) PH_3 (phosphine)
*C) H_2O (water)
D) H_2Se (hydrogen selenide)

37) The solid with the **lowest** solubility in water is a(n)

A) ionic solid
B) non polar molecular solid
C) polar molecular solid
*D) network solid

38) The substance listed below that has non-polar molecules, covalent bonding between atoms in the molecule and only van der Waals' forces acting between the molecules is

A) KCl (potassium chloride)
*B) CCl_4 (carbon tetrachloride)
C) HBr (hydrogen bromide)
D) NH_3 (ammonia)

39) The **most polar** compound in the following list is

*A) solid KCl (potassium chloride)
B) liquid CH_4 (methane)
C) liquid HBr (hydrogen bromide)
D) gaseous Ne (neon)

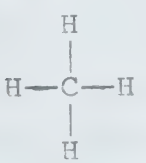
40) Consider the following data for five substances:

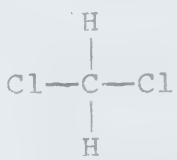
SUBSTANCE	MELTING POINT (^{o}C)	BOILING POINT (^{o}C)	ELECTRICAL CONDUCTIVITY	
			Solid	Liquid
I	776	1500	poor	good
II	-39	356	good	good
III	-190	-42	poor	poor
IV	961	1950	good	good
V	1420	2355	poor	good

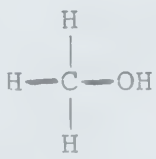
The substances that may involve metallic bonding are

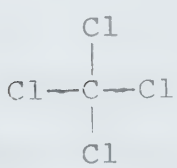
A) I and V
*B) II and IV
C) II and V
D) I and IV

41) Which of the following compound(s) when mixed with water show great attraction for water molecules and therefore high solubility in water?

I) 

II) 

III) 

IV) 

V)

A) I and IV only
B) II and IV only
*C) III and V only
D) V only

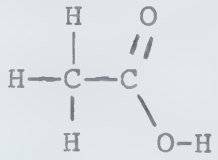
13A.11.

42) The main attractive forces between most types of polar molecules are called

*A) dipole-dipole attractions
B) hydrogen bonding
C) van der Waals attractions
D) ionic bonding

43) The molecule listed below that has the least polar character is

I)



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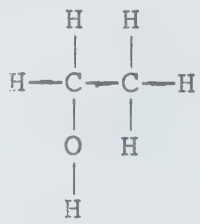
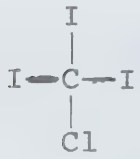


III)



IV)

A) I
B) II
C) III
*D) IV

44) The compound listed below that is held together in the solid form by the weakest bonds is

A) CsCl
*B) CO_2
C) CCl_4
D) H_2O

13A.11.

45) A good solvent for polar molecules such as HCl (hydrogen chloride) is

*A) H_2O (water)
B) C_6H_6 (benzene)
C) CCl_4 (carbon tetrachloride)
D) CS_2 (carbon disulphide)

46) A transparent crystalline substance is found to be rather soft and to have a low melting point. The crystal is most probably an example of a(n)

*A) molecular solid
B) ionic solid
C) network solid
D) metallic solid

47) The substance in which all the bonds are only van der Waals forces is

A) N_2 (nitrogen)
B) K (potassium)
*C) Ne (neon)
D) H_2 (hydrogen)

13A.11.

48) Consider the following data for the substances listed:

SUBSTANCE	MELTING POINT (°C)	BOILING POINT (°C)	ELECTRICAL CONDUCTIVITY	
			Solid	Liquid
I	776	1500	poor	good
II	-39	356	good	good
III	-190	-42	poor	poor
IV	961	1950	good	good

The substance that is a non-polar molecular crystal is

A) I
B) II
*C) III
D) IV

49) Consider the following data for the substances listed:

SUBSTANCE	MELTING POINT (°C)	BOILING POINT (°C)	ELECTRICAL CONDUCTIVITY	
			Solid	Liquid
I	776	1500	poor	good
II	-39	356	good	good
III	-190	-42	poor	poor
IV	1420	2355	poor	poor

The substance that is an ionic solid at STP is

*A) I
B) II
C) III
D) IV

50) A good example of a 3-dimensional covalent network solid is

A) NaCl (sodium chloride)
B) CH_4 (methane)
*C) SiC (silicon carbide)
D) Fe (iron)

13A.11.

51) The compound in the following list that is composed of polar covalent molecules is

A) CsCl

B) CO_2

C) CCl_4

*D) H_2O

52) Consider the following information about four substances:

Substance	Melting Point (^{o}C)	Electrical Conductivity in solid	Electrical Conductivity in liquid	Water solubility
I	1050	low	high	high
II	75	low	low	low
III	1234	high	high	low
IV	1913	low	low	low

The substance which is most likely to be a metal is

A) I

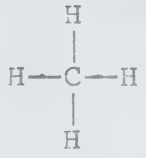
B) II

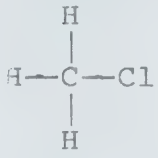
*C) III

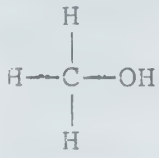
D) IV

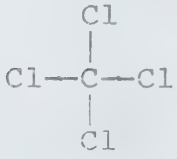
13A.11.

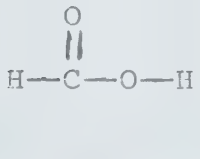
53) Which of the following molecules would be polar?

I) 

II) 

III) 

IV) 

V) 

A) I and IV only
B) II and IV only
C) all of them
*D) II, III and V only

54) Metal crystals tend to be hard but not brittle, whereas ionic crystals tend to be hard and brittle. Which is best model to explain this difference?

A) There are no ions in metallic crystals.
B) There are fewer electrons per atom in metals than in ionic aggregates.
C) Atoms are less tightly packed in ionic crystals than in metals, resulting in decreased coordination numbers in ionic crystals.
*D) The attractive forces in metallic crystals are relatively unaffected by shifts in layers of atoms, whereas attractive forces in ionic crystals are greatly affected if layers shift.

55) Water and oil do not mix because

A) their densities are different
B) oil is a polar liquid
C) water is a non-polar liquid
*D) no bonds form between the oil and the water

13A.11.

56) 2270 J of heat is required to vaporize one gram of H_2O (water) while only 590 J is needed to vaporize one gram of C_8H_{18} (octane). The reason why the heat of vaporization of H_2O is much higher than that of C_8H_{18} is that

A) octane has polar molecules
B) octane gives off little energy when it vaporizes
C) molecules of water are not as massive as molecules of octane
*D) the intermolecular forces of attraction are stronger in water than in octane

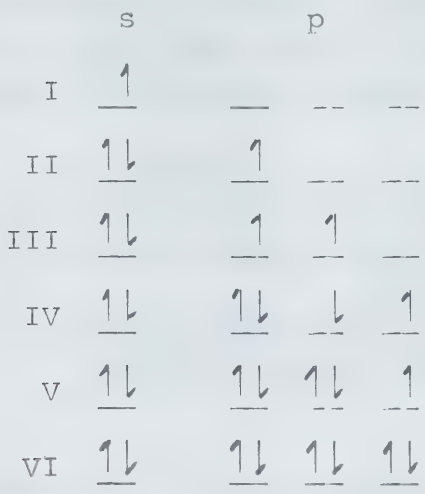
57) Consider the properties for the substances listed:

SUBSTANCE	MELTING POINT (K)	BOILING POINT (K)	ELECTRICAL CONDUCTIVITY OF SOLID	ELECTRICAL CONDUCTIVITY OF LIQUID	SOLUBILITY IN WATER
1	1050	1770	poor	good	soluble
2	234	629	good	good	low solubility
3	83	231	poor	poor	low solubility
4	1230	2220	good	good	low solubility
5	1690	2630	poor	poor	low solubility

In which one of these substances are the primary intermolecular forces van der Waals forces?

A) 1
B) 2
*C) 3
D) 4

58) Which element would be most likely to form a network solid?



A) I
B) V
*C) III
D) IV

59) Germanium, Ge, is a solid whose atoms are covalently bonded to each other, much like those of carbon atoms in the diamond crystal.
Potassium chloride, KCl, is a white crystalline solid.
Methane, CH_4, is a gas which can be liquefied under high pressures and low temperatures.
Rubidium metal, Rb, is very malleable and is an excellent conductor of electricity.
Which substance in the solid state consists of particles held together by van der Waals forces only?

A) Ge
B) KCl
*C) CH_4
D) Rb

60) Glycerine, $C_3H_5(OH)_3$, is an alcohol.
Potassium chloride, KCl, is a white crystalline solid.
Methane, CH_4, is a gas which can be liquefied under high pressures and at low temperatures.
Rubidium metal, Rb, is very malleable and is an excellent conductor of electricity.
Which substance would have the lowest heat of vaporization?

A) $C_3H_5(OH)_3$
B) KCl
*C) CH_4
D) Rb

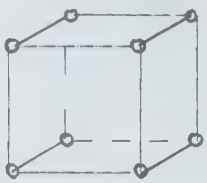
61) From the following table one could predict that carbon tetrachloride, CCl_4, would be a good solvent for

SUBSTANCE		m.p. (K)	b.p. (K)
He	(helium)	3.3	4.1
NaCl	(sodium chloride)	1074.4	1686
CCl_4	(carbon tetrachloride)	250	349.8
SiO_2	(silicon dioxide)	above 2000	about 2500
C	(diamond)	above 3800	about 4500
CH_4	(methane)	89	111
H_2O	(water)	273	373
KOH	(potassium hydroxide)	633.4	about 1700
HF	(hydrogen fluoride)	189.9	292.54

A) SiO_2 (silicon dioxide)
B) C (diamond)
*C) CH_4 (methane)
D) KOH (potassium hydroxide)

13A.11.

62) The number of **nearest neighbours** (coordination number) for the atom of the crystal represented in the following diagram is



*A) 6
B) 8
C) 12
D) 4

63) The bonding in granite is most likely to be

A) ionic
*B) covalent network
C) metallic
D) van der Waals

13A.11.

64) In which of the substances in the table does hydrogen bonding occur?

SUBSTANCE		m.p. (K)	b.p. (K)
He	(helium)	3.3	4.1
NaCl	(sodium chloride)	1074.4	1686
CCl_4	(carbon tetrachloride)	250	349.8
SiO_2	(silicon dioxide)	above 2000	about 2500
C	(diamond)	above 3800	about 4500
CH_4	(methane)	89	111
H_2O	(water)	273	373
KOH	(potassium hydroxide)	633.4	about 1700
HF	(hydrogen fluoride)	189.9	292.54

A) He, CCl_4 and CH_4

*B) H_2O and HF

C) NaCl and KOH

D) C and SiO_2

65) Which of the substances in the table are examples of network solids?

SUBSTANCE		m.p. (K)	b.p. (K)
He	(helium)	3.3	4.1
NaCl	(sodium chloride)	1074.4	1686
CCl_4	(carbon tetrachloride)	250	349.8
SiO_2	(silicon dioxide)	above 2000	about 2500
C	(diamond)	above 3800	about 4500
CH_4	(methane)	89	111
H_2O	(water)	273	373
KOH	(potassium hydroxide)	633.4	about 1700
HF	(hydrogen fluoride)	189.9	292.54

A) He, CCl_4 and CH_4

B) H_2O and HF

C) NaCl and KOH

*D) C and SiO_2

66) Glycerine, $C_3H_5(OH)_3$, is an alcohol.
Potassium chloride, KCl, is a white crystalline solid.
Methane, CH_4, is a gas which can be liquefied under high pressures and at low temperatures.
Rubidium metal, Rb, is very malleable and is an excellent conductor of electricity.
In which substance would hydrogen bonding be most likely to occur?

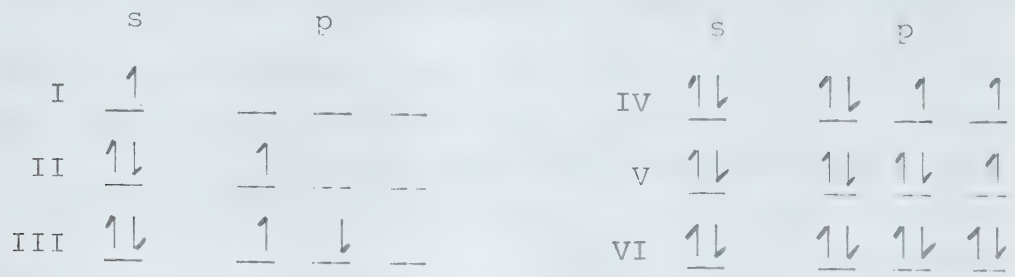
*A) $C_3H_5(OH)_3$

B) KCl

C) CH_4

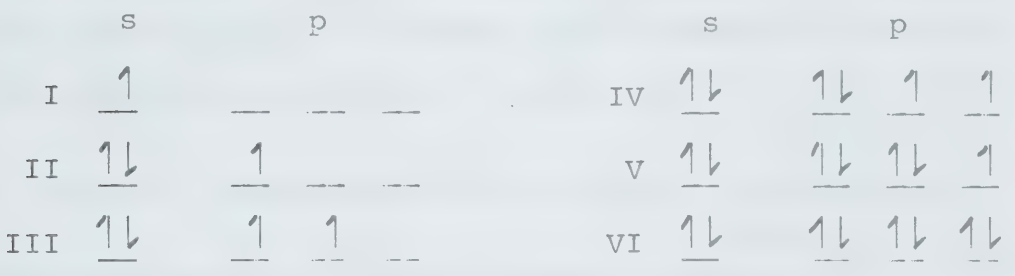
D) Rb

67) Which of the elements represented in the diagram would most likely be metallic in character?



A) I and III
B) III, IV and V
C) II and III
*D) I and II

68) Which of the following pairs of the elements represented in the diagram would most likely form ionic bonds when reacting together?



A) I and II
*B) I and V
C) III and IV
D) IV and V

13A.11.

69) Germanium, Ge, is a solid whose atoms are covalently bonded to each other, much like those of carbon atoms in the diamond crystal.
Potassium chloride, KCl, is a white crystalline solid.
Methane, CH_4, is a gas which can be liquefied only under high pressures and at low temperatures.
Rubidium metal, Rb, is very malleable and is an excellent conductor of electricity.
Which substance would be the hardest in the solid state?

*A) Ge
B) KCl
C) CH_4
D) Rb

70) The boiling point of water, $H_2O(\ell)$, is considerably higher than that of hydrogen sulphide, $H_2S(\ell)$. Which of the following statements best accounts for this difference?

A) Van der Waals forces are greater in $H_2O(\ell)$ than in $H_2S(\ell)$.
*B) Hydrogen bonding is more pronounced in $H_2O(\ell)$ than in $H_2S(\ell)$.
C) The molecular mass of H_2O is different from that of H_2S.
D) H_2O molecules have more unshared pairs of valence electrons than do molecules of H_2S.

71) The micas, minerals composed largely of silicon and oxygen, are unique in that they can be separated easily into very thin sheets which are nearly transparent and resistant to heat. The type of structure found in micas can best be described as

A) one-dimensional network solid
*B) two-dimensional network solid
C) three-dimensional network solid
D) ionic solid

72) A white crystalline substance melts below 100°C. It does not conduct electricity either as a solid or when melted. This substance is most likely to be

A) a pure metal
B) an ionic compound
C) a covalent network solid
*D) a molecular solid

73) Germanium, Ge, is a solid whose atoms are covalently bonded to each other, much like those of carbon atoms in the diamond crystal.
Glycerine, $C_3H_5(OH)_3$, is an alcohol.
Potassium chloride, KCl, is a white crystalline solid.
Rubidium metal, Rb, is very malleable and is an excellent conductor of electricity.
Which substance would have the highest melting point?

*A) Ge
B) $C_3H_5(OH)_3$
C) KCl
D) Rb

74) All chemical bonds result from

A) one or more electrons lost by certain atoms and gained by others
*B) one or more electrons being simultaneously attracted by two or more nuclei
C) overlapping of two electron-filled orbitals having different energy levels
D) overlapping of unoccupied orbitals of two or more atoms

75) Which of the substances in the table, when in solid state, are molecular solids composed of polar molecules?

SUBSTANCE		m.p. (K)	b.p. (K)
He	(helium)	3.3	4.1
NaCl	(sodium chloride)	1074.4	1686
CCl_4	(carbon tetrachloride)	250	349.8
SiO_2	(silicon dioxide)	above 2000	about 2500
C	(diamond)	above 3800	about 4500
CH_4	(methane)	89	111
H_2O	(water)	273	373
KOH	(potassium hydroxide)	633.4	about 1700
HF	(hydrogen fluoride)	189.9	292.54

A) He, CCl_4 and CH_4

*B) H_2O and HF

C) NaCl and KOH

D) C and SiO_2

76) Which of the substances in the table form network solids held together by bonds which are predominately covalent.

SUBSTANCE		m.p. (K)	b.p. (K)
He	(helium)	3.3	4.1
NaCl	(sodium chloride)	1074.4	1686
CCl_4	(carbon tetrachloride)	250	349.8
SiO_2	(silicon dioxide)	above 2000	about 2500
C	(diamond)	above 3800	about 4500
CH_4	(methane)	89	111
H_2O	(water)	273	373
KOH	(potassium hydroxide)	633.4	about 1700
HF	(hydrogen fluoride)	189.9	292.54

A) H_2O and HF

B) NaCl and KOH

*C) C and SiO_2

D) He and CCl_4

77) In polyatomic molecules, molecular polarity results from

A) equal sharing of electrons between atoms

*B) bond dipoles which do not cancel each other

C) electron promotion

B) orbital hybridization

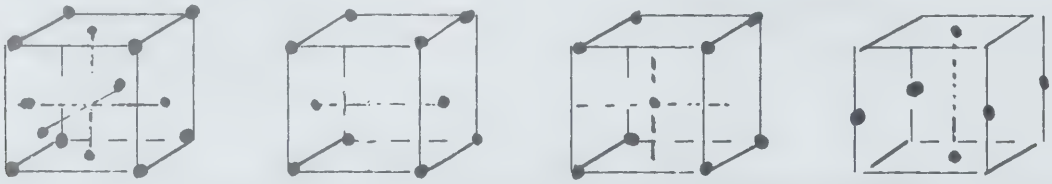
78) Intermolecular forces in C_xH_y (organic oil-like compounds) are mainly due to

A) interionic attraction
B) hydrogen bonds
*C) van der Waals forces
D) dipole dipole bonds

79) A certain solid consists of ions bound in a crystal lattice. Which of the following properties would be expected to be exhibited by the solid?

A) low melting point
*B) hard and brittle
C) good conductor
D) low solubility in polar solvents.

80) Which of the following figures best represents cubic closest (face-centred) packing?



I II III IV

*A) I
B) II
C) III
D) IV

13A.11.

81) A chemical bond forms between two atoms.

*A) when the attractive forces between the atoms are greater than the repulsive forces

B) because the molecule formed is less stable than the individual atoms

C) only when equal sharing of the bonding electrons results

D) because an increase of energy results from such bond formation

82) Which of the following bonds has the least ionic character?

A) H-Cl

B) Br-Cl

C) S-Cl

*D) Cl-Cl

83) The sharing of electrons in bond formation always involves

A) formation of positive and negative ions

B) formation of polar molecules

C) shared electrons being attracted more by one atom than another

*D) lower energy content for the bonded than the unbonded atoms

84) Metals are generally better conductors of both heat and electricity than non-metals because

A) metals have higher melting points

*B) valence electrons in metals are more mobile

C) metals contain more electrons

D) metals have higher ionization energies

85) Which of the following compounds is the best example of a network solid?

A) $NaCl(s)$ (table salt)

B) $C_{25}H_{52}(s)$ (paraffin)

C) $S_8(s)$ (sulphur)

*D) $SiO_2(s)$ (quartz)

86) The strength of intermolecular "bonds" due to van der Waals forces increases in a given family (such as the halogens) as the

*A) number of electrons in the atom increases

B) molecular size decreases

C) temperature increases

D) molecules become more nearly spherical

13A.12.

1) An analysis of a compound gives the following composition by mass:

C: 40%, H: 6.7%, O: 53%

The simplest formula of the compound is

A) C_2HO

B) CHO_2

*C) CH_2O

D) C_2HO_2

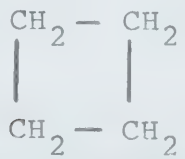
2) Which two of the following hydrocarbons are structural isomers?

I $CH_3CH_2CH_2CH_3$

II $CH_3CH = CHCH_3$

III $CH_3C \equiv CCH_3$

IV



A) I and III

*B) II and IV

C) II and III

D) III and IV

3) Hexane (C_6H_{14}) is burned in excess oxygen gas. The equation representing the combustion of the hexane is

*A) $C_6H_{14} + \frac{19}{2} O_2 \rightarrow 6CO_2 + 7H_2O$

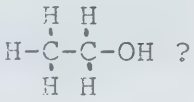
B) $C_6H_{14} + 19\ O_2 \rightarrow 6CO_2 + 7H_2O$

C) $C_6H_{14} + \frac{13}{2} O_2 \rightarrow 6CO + 7H_2O$

D) $C_6H_{14} + 13\ O_2 \rightarrow 6CO + 7H_2O$

13A.12.

4) Which type of organic compound is represented by the structural formula,



A) an aldehyde
B) an amine
*C) an alcohol
D) a ketone

5) The formula for methyl acetate is

A) $CH_3CH_2COCH_3$
*B) CH_3COOCH_3
C) $CH_3COOCH_2CH_3$
D) $CH_3CHOHCH_3$

6) When an alkene is hydrated, the product is a(an)

A) acid
B) aldehyde
C) ketone
*D) alcohol

7) A certain gas contains 85.6% carbon and 14.38% hydrogen by mass. The simplest formula of the gas is

A) CH
*B) CH_2
C) C_2H
D) CH_4

13A.12.

8) In the reaction for making the ester, ethyl acetate, 2.5 mol of acetic acid was mixed with 3.2 mol of ethanol and 1.8 mol of the ester was obtained. The percentage yield of the ester was

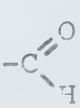
A) 30%
B) 56%
*C) 72%
D) 78%

9) Which of the following compounds may be a **structural isomer** of ethyl alcohol?

A) CH_3CHO
B) CH_3OH
*C) CH_3OCH_3
D) CH_3CH_3

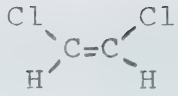
10) A gas which has a relative molecular mass 42.0 has the simplest (empirical) formula CH_2. The molecular formula of the gas is

A) C_2H_4
*B) C_3H_6
C) C_4H_8
D) C_6H_{12}

11) The functional group  is found in all

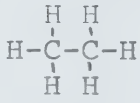
A) ketones
B) alcohols
*C) aldehydes
D) acids

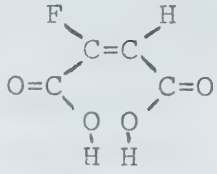
13A.12.

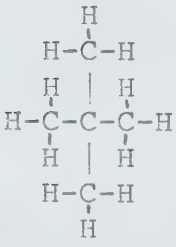
12) The correct name for the compound  is

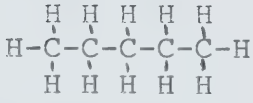
*A) cis-dichloroethene
B) trans-dichloroethene
C) 1,2-dichloroethane
D) 1,1-dichloroethene

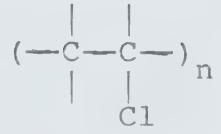
13) Which one of the following substances can form cis-trans isomers

A) 

* C) 

B) 

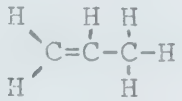
D) 

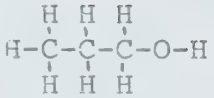
14) Polyvinyl chloride  is a(n)

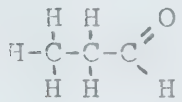
*A) addition polymer
B) condensation polymer
C) polyamide
D) polyester

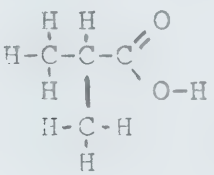
13A.12.

15) The structural diagram for propanol is

I) 

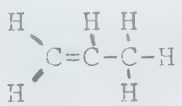
III) 

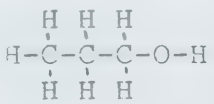
II) 

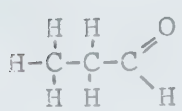
IV) 

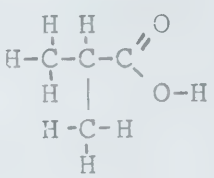
A) I
B) II
*C) III
D) IV

16) Which pair of the compounds could combine in the presence of concentrated sulphuric acid, to produce an ester?

I) 

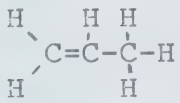
III) 

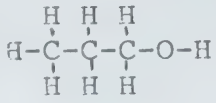
II) 

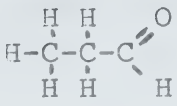
IV) 

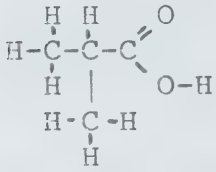
A) I and II
B) II and III
C) II and IV
*D) III and IV

17) Which formula represents an unsaturated hydrocarbon?

I) 

III) 

II) 

IV) 

*A) I
B) II
C) III
D) IV

18) The coefficient of oxygen when the following equation is balanced for one mole of C_6H_6 (benzene) is

$$\underline{1}\ C_6H_6(\ell) + O_2(g) \rightarrow CO_2(g) + H_2O(\ell)$$

A) 7
*B) 15/2
C) 9
D) 15

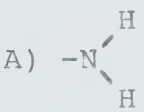
19) The chemical reaction between an organic acid and an alcohol is known as

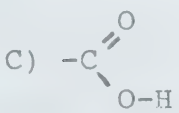
A) saponification
B) hydrogenation
C) hydrolysis
*D) esterification

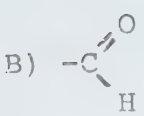
20) The alkyl radical, C_6H_{13}, is called .

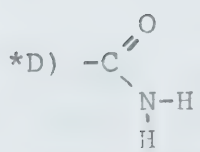
A) heptyl
*B) hexyl
C) heptanyl
D) hexanyl

21) The structural diagram for the amide functional group is

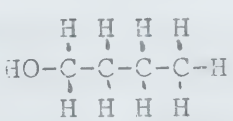
A) 

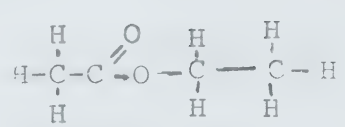
C) 

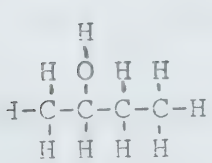
B) 

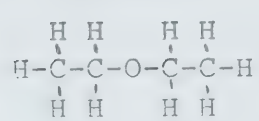
*D) 

22) Consider the following structural formulae:

I) 

III) 

II) 

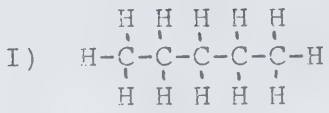
IV) 

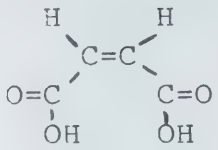
Which compound might be oxidized most readily, to make ethyl methyl ketone?

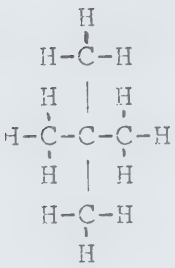
A) I
*B) II
C) III
D) IV

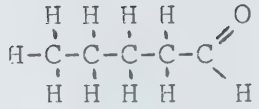
13A.12.

23) Which two of the following substances are structural isomers?

I) 

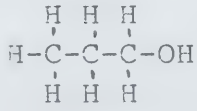
III) 

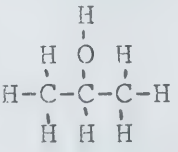
II) 

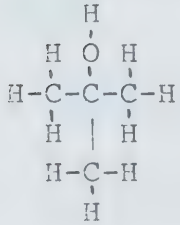
IV) 

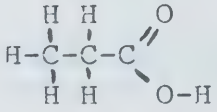
*A) I and II
B) III and IV
C) II and III
D) I and IV

24) The structural formula that represents a **secondary alcohol** is

A) 

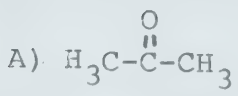
*C) 

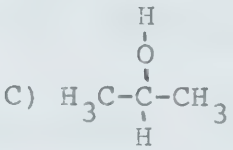
B) 

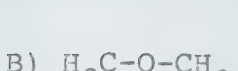
D) 

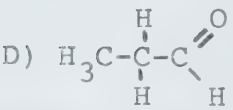
13A.12.

25) Which of the following compounds is an ether?

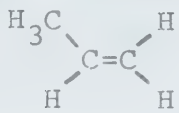
A) 

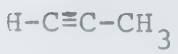
C) 

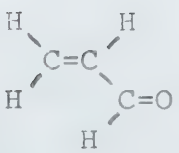
B) 

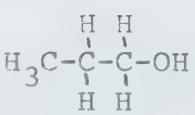
D) 

26) A solution of bromine water is added to one mole of each of the compounds

I) 

III) 

II) 

IV) 

Which of the compounds (per mole) will decolorize the largest volume of bromine water?

A) I
B) II
*C) III
D) IV

27) Organic acids are generally

A) soluble in water
B) strong electrolytes
C) strong acids
*D) weak acids

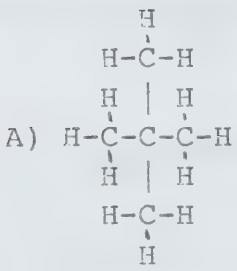
13A.12.

28) A student is given 4 bottles each containing a different compound. The following information concerning melting temperature, boiling temperature and physical state at room temperature is given for the compound in each bottle.

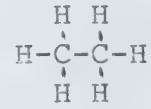
Bottle Label	Melting ^{o}C Temperature	Boiling ^{o}C Temperature	State at Room Temperature
P	-183	-89	gas
Q	-20	9	gas
S	-78.5	138	liquid
T	284	290	solid

Which of the compounds is contained in bottle T?

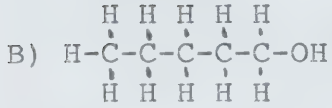
A)



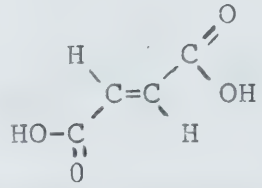
C)



B)

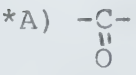


*D)

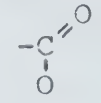


29) The functional group of a ketone is

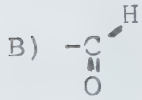
*A)



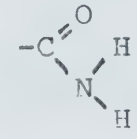
C)



B)



D)

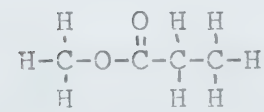


30) Polypropylene is commonly used for ropes and rugs. It is produced by the polymerization of

A) propane
*B) propene
C) propyne
D) cyclopropane

31) Which of the following hydrocarbons causes a bromine solution to become colourless?

A) C_2H_6 (ethane)
*B) C_2H_4 (ethene)
C) C_6H_6 (benzene)
D) $C_2H_4Br_2$ (dibromoethane)

32) The ester  is called

A) ethyl formate
B) methyl acetate
C) propyl formate
*D) methyl propanoate

33) The model that carbon forms four sp^3 orbitals is used to explain the fact that

A) carbon compounds are generally insoluble in water
*B) the four bonds in methane (CH_4) are identical
C) acetic acid, CH_3COOH, is a weak acid
D) graphite and diamond are allotropes

34) Which of the following statements applies to unsaturated hydrocarbons?

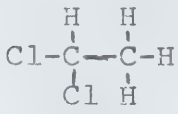
*A) The names in the compounds containing double bonds end in "ene".

B) The prefix "butyl" denotes three carbon atoms.

C) They are quite unreactive.

D) The most characteristic type of reaction is substitution.

35) The structural formula



represents

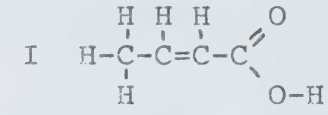
A) 1,1, dichloromethane

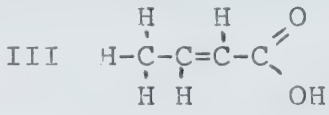
B) 1,2, dichloromethane

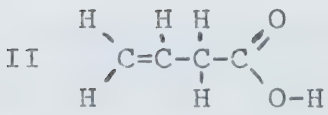
C) 1,2, dichloroethane

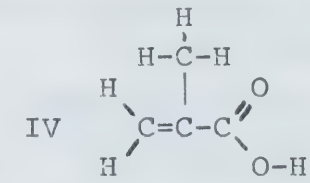
*D) 1,1, dichloroethane

36) Consider these four possible structures with the formula C_4H_6O:









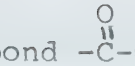
Cis-trans isomerism is illustrated by

A) I and II

*B) I and III

C) I and IV

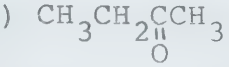
D) II and III

37) The bond  can be

*A) an ester linkage
B) a peptide bond
C) an amide linkage
D) an amino acid linkage

38) Which of the following formulae is representative of a ketone?

A) CH_3CH_2CHO

*B) 

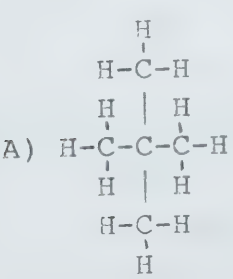
C) CH_3CH_2COOH

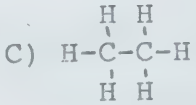
D) CH_3OCH_3

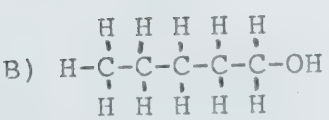
39) A student is given 4 bottles each containing a different compound. The following information concerning melting temperature, boiling temperature and physical state at room temperature is given for the compound in each bottle.

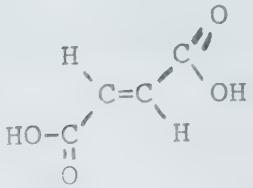
Bottle Label	Melting ^{o}C Temperature	Boiling ^{o}C Temperature	State at Room Temperature
P	-183	-89	gas
Q	-20	9	gas
S	-78.5	138	liquid
T	284	290	solid

Which of the following compounds is contained in bottle S?

A) 

C) 

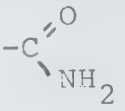
*B) 

D) 

13A.12.

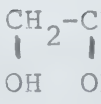
40) When butanoic (butyric) acid reacts with ethanol the product is

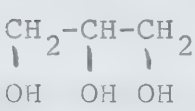
*A) ethyl butanoate
B) isobutyl acetate
C) ethyl propionate
D) butyl ethanoate

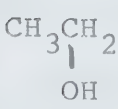
41) The structural formula  is representative of an

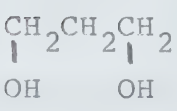
A) amine
B) alcohol
*C) amide
D) amino benzoate

42) Which of the following is the formula for glycerol?

A) 

*B) 

C) 

D) 

43) Soap shaken with a mixture of water and fuel oil produces a(n)

A) pure substance
*B) emulsion
C) solution
D) colloidal dispersion

13A.12.

44) The equation

is an example of a(n)

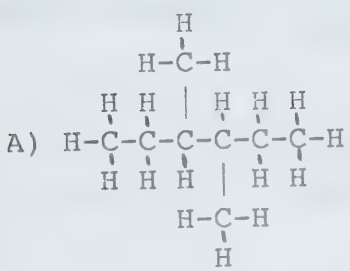
A) esterification reaction
B) neutralization reaction
*C) hydrolysis reaction
D) acidification reaction

45) The bond formed between amino acids in a protein is called a

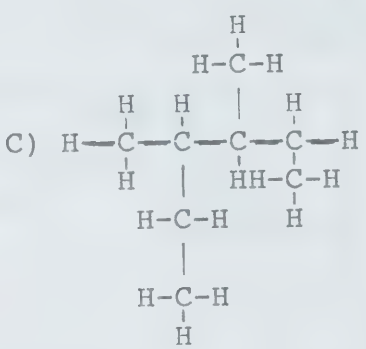
A) peptine bond
*B) peptide bond
C) polypeptide bond
D) ketyl bond

46) The formula of 2,3,4-trimethylpentane is

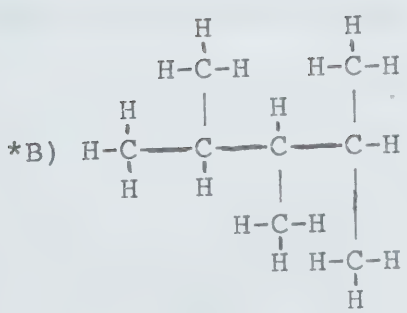
A)



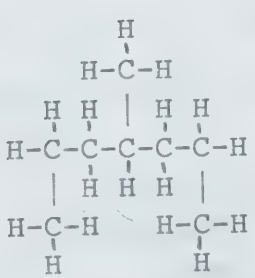
C)



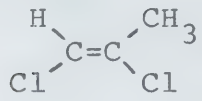
*B)



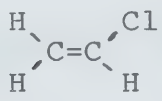
D)



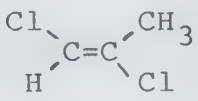
47) Which of the following are cis-trans isomers?



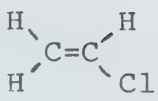
I



II



III



IV

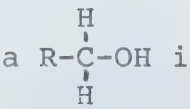
*A) I and III
B) III and IV
C) II and IV
D) II and III

48) is the functional group of

A) acids
B) ketones
C) amines
*D) aldehydes

49) A hydrocarbon consists of 86% carbon and has a relative molecular mass of 42. The compound is

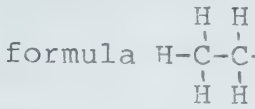
A) cyclopropene
B) propyne
*C) propene
D) propane

50) The type of organic compound represented by the general formula  is called a(n)

*A) alcohol
B) aldehyde
C) amine
D) ketone

13A.12.

51) Which type of organic compound is represented by the formula



A) an aldehyde
B) an amine
*C) an alcohol
D) an acid

52) How many structural isomers have the molecular formula C_5H_{12}?

A) 5
B) 2
*C) 3
D) 4

53) In the preparation of an ester by the action of an acid on an alcohol, concentrated H_2SO_4 (sulphuric acid) is also added in order to

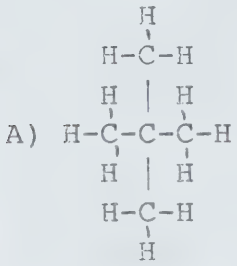
A) prevent the reaction taking place too fast
*B) act as a catalyst
C) prevent hydrolysis of the ester
D) act as an oxidising agent

54) A student is given 4 bottles each containing a different compound. The following information concerning melting temperature, boiling temperature and physical state at room temperature is given for the compound in each bottle.

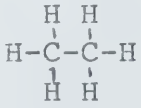
Bottle Label	Melting Temperature ^{o}C	Boiling Temperature ^{o}C	State at Room Temperature
P	-183	-89	gas
Q	-20	9	gas
S	-78.5	138	liquid
T	284	290	solid

Which of the compounds is contained in bottle P?

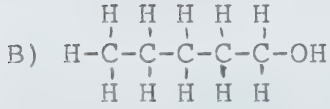
A)



*C)



B)



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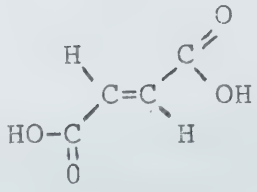


55) The general formula for a **cycloalkane** is

A) C_nH_n

B) C_nH_{2n+2}

*C) C_nH_{2n}

D) C_nH_{2n-2}

56) Natural gas is mainly

*A) CH_4 (methane)

B) C_2H_6 (ethane)

C) C_3H_8 (propane)

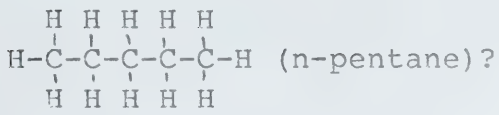
D) C_4H_{10} (butane)

13A.12.

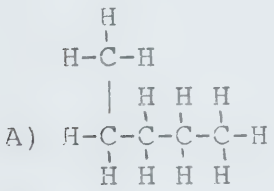
57) Acetylene (C_2H_2)

A) has a relative molecular mass of 13
*B) is a linear molecule
C) has a bond angle of 120^{o}
D) burns to produce CH_4 (methane)

58) Which of the following is a structural isomer of



(n-pentane)?

A)


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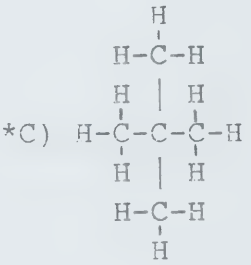

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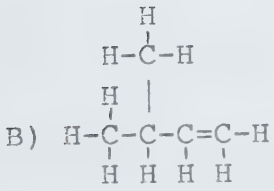

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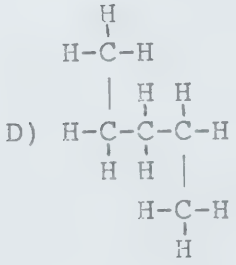
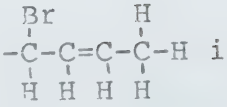
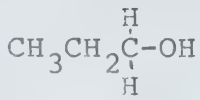


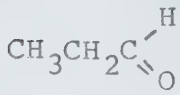
59) The name of the compound

is

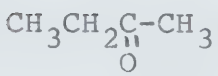
*A) 1-bromo-2-butene
B) 4-bromo-2-butyne
C) 4 bromo-1-methyl-1-propyne
D) 1-bromo-2-butyne

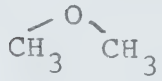
13A.12.

60) Which of the following formulae is representative of a ketone?

A) 

B) 

*C) 

D) 

13A.14.

1) A solution that turns blue when potassium ferricyanide ($K_3Fe(CN)_6$) is added must contain

*A) $Fe^{2+}(aq)$

B) $Fe^{3+}(aq)$

C) $NO_3^-(aq)$

D) $Ag^+(aq)$

13A.15.

1) Heavy water is

A) an allotrope of water
*B) a form of water containing deuterium
C) water at the temperature of its maximum density, 277 K
D) a substance consisting of H_3O molecules

2) A radioactive element A with mass number 214 and atomic number 84 emits an alpha particle and changes to element B. Element B emits a beta particle and is converted to element C. Elements B and C are represented by

*A) $^{210}_{82}B$ and $^{210}_{83}C$

B) $^{212}_{82}B$ and $^{210}_{81}C$

C) $^{212}_{80}B$ and $^{211}_{81}C$

D) $^{210}_{80}B$ and $^{210}_{81}C$

3) Bi, atomic number 83, emits a beta-particle; which of the following statements is/are correct?

I Lead, atomic number 82, is formed
II Polonium, atomic number 84, is formed
III The mass number increases
IV The mass number decreases

A) I and IV
B) II and III
C) I only
*D) II only

4) Consider the nuclear reaction

$$^{245}_{95}Am + \underline{\quad ? \quad} \rightarrow {}^{247}_{97}Bk + 2\ {}^{1}_{0}n$$

Which of the following is necessary to complete the equation?

*A) $^{4}_{2}He$

B) $2\ {}^{1}_{1}H$

C) $4\ {}^{1}_{1}H$

D) $4\ {}^{0}_{-1}e$

5) In the fission of uranium the following reaction occurs:

$$^{235}_{92}U + {}^{1}_{0}n \rightarrow {}^{141}_{56}Ba + 3\ {}^{1}_{0}n + \text{———}$$

The missing product is

A) $^{94}_{36}Kr$

*B) $^{92}_{36}Kr$

C) $^{90}_{35}Br$

D) $^{88}_{39}Y$

6) One of the isotopes of plutonium, $^{242}_{94}Pu$, when bombarded by alpha particles forms a new element and two neutrons according to the following equation:

$$^{242}_{94}Pu + ^{4}_{2}He \rightarrow \text{(new element)} + 2\ ^{1}_{0}n$$

The new element, X, is

A) $^{238}_{96}X$

B) $^{238}_{92}X$

C) $^{244}_{92}X$

*D) $^{244}_{96}X$

7) Which of the following statements are correct?

I The neutron to proton ratio is reduced by the emission of a beta-particle

II It is reasonable to predict that the $^{16}_{8}O$ isotope would be stable

III Tracers are radioactive isotopes that can be used to determine reaction mechanisms

IV Sr-90 with a half-life of 28 years is more stable than Co-60 with a half-life of 5.3 years.

A) II and IV

B) II, III and IV

C) I and II

*D) All are correct

8) $^{60}_{27}Co$, a radioisotope, is used to destroy cancer cells. The equation which best illustrates this use is

A) $^{60}_{27}Co + {}^{1}_{0}n \rightarrow {}^{61}_{27}Co$

B) $^{60}_{27}Co + {}^{0}_{-1}e \rightarrow {}^{60}_{26}Fe + {}^{0}_{-1}e$

*C) $^{60}_{27}Co \rightarrow {}^{60}_{28}Ni + {}^{0}_{-1}e$

D) $^{60}_{27}Co + {}^{4}_{2}He \rightarrow {}^{62}_{29}Cu + 2\ {}^{1}_{0}n$

9) The term radioactivity refers to

A) activity of electronic devices
*B) spontaneous disintegration of certain atomic nuclei
C) loss of electrons from the outer orbits of the atom
D) reactions between elements catalyzed by radio waves

10) A radioactive isotope is one

A) which has unpredictable properties
B) which when heated gives off radio waves
C) whose nucleus undergoes a change only when energy is absorbed
*D) whose nucleus spontaneously changes, liberating energy

11) A property of an alpha particle is that it

A) is a hydrogen nucleus
B) has 4 protons in its nucleus
*C) consists of 2 protons and 2 neutrons
D) has a charge of +1

13A.15.

12) A 36.0 g sample of a radioactive isotope "X" decays for three half-lives. The mass of "X" remaining is

*A) 4.50 g
B) 9.00 g
C) 3.00 g
D) 18.00 g

13) If radium, atomic number 88, emits an alpha-particle from the nucleus, which one of the following elements is formed?

A) polonium atomic number 84
*B) radon atomic number 86
C) francium atomic number 87
D) thorium atomic number 90

14) When an atom emits a beta-particle from its nucleus the resultant nucleus contains

A) one less proton and one more neutron
B) one more proton and one more neutron
*C) one more proton and one less neutron
D) one less proton and the same number of neutrons

15) Elements are radioactive if

*A) their atomic nuclei are unstable
B) the electrons outside the nuclei are easily lost
C) their atoms gain electrons readily
D) they are chemically active

13A.15.

16) Of a 60.0 g sample of a radioactive isotope what mass will remain at the end of four half-lives?

*A) 3.75 g
B) 7.50 g
C) 15.0 g
D) 960 g

17) What particle is required to complete the following nuclear equation?

$$^{209}_{83}Bi + ^{2}_{1}H \rightarrow ^{210}_{84}Po + \text{-----}$$

A) proton
B) beta particle
*C) neutron
D) alpha particle

18) $^{235}_{92}U$ (uranium 235) emits an alpha particle producing a radioactive substance (X). This substance (X) disintegrates emitting a beta-particle and producing another element (Y). Substance Y is

A) $^{232}_{89}Ac$

B) $^{233}_{91}Pa$

C) $^{237}_{93}Np$

*D) $^{231}_{91}Pa$

13A.15.

19) Which of the following statements are correct?

I Pa-234 with a half-life of 6.7 h is less stable than Fe-55 with a half-life of 2.7 years.

II The neutron to proton ratio is decreased by the emission of a beta-particle.

III Artificial isotopes are made by using radioactive isotopes as tracers.

IV It would be reasonable to predict that $^{34}_{15}P$ would be unstable.

A) II and III
B) I and III
*C) I, II and IV
D) I and IV

20) Radioactive carbon $^{14}_{6}C$ has a half-life of 5700 years. It changes slowly to ordinary nitrogen $^{14}_{7}N$ according to the equation:

$$^{14}_{6}C \rightarrow {}^{14}_{7}N + (?)$$

The type of radioactivity illustrated here is

A) alpha-particle decay
*B) beta-particle decay
C) K-electron capture
D) positron emission

21) All of the elements in the seventh row of the Periodic Table are

A) found in natural uranium
B) man-made
C) products of nuclear fission
*D) radioactive

13A.15.

22) The equation, $E = mc^2$, predicts to the quantity of energy

A) required to start a nuclear reaction
*B) generated during a nuclear reaction
C) necessary to moderate a nuclear reactor
C) involved in producing radio-isotopes

23) The following nuclear reaction

$$^{131}_{53}I \rightarrow ^{131}_{54}Xe + ^{0}_{-1}e$$

is an example of

A) alpha decay
*B) beta decay
C) positron decay
D) fission

24) A radioactive isotope P has a half life of 8 h, and an isotope Q has a half life of 12 h. A mixture is prepared containing equal masses of P and Q. What is the mass ratio of P to Q remaining in the mixture after twenty-four (24) h?

A) 1:1
*B) 1:2
C) 2:1
D) 2:3

13A.15.

25) Consider the following equation:

$$^{235}_{92}U + 1\ ^{1}_{0}n \rightarrow\ ^{147}_{59}Pr + 3\ ^{1}_{0}n + \text{———}?$$

The missing product is an isotope of

*A) arsenic
B) rubidium
C) hydrogen
D) uranium

26) Control rods of cadmium are used in a nuclear reactor to

A) absorb protons
B) supply protons
C) speed up neutrons
*D) slow down neutrons

27) When $^{204}_{82}Pb$ decays to element Y by emitting an alpha particle, the new product is

A) $^{202}_{82}Pb$

B) $^{200}_{78}Pt$

*C) $^{200}_{80}Hg$

D) $^{208}_{88}Ra$

28) When fission of a heavy nucleus has taken place, and new neutrons are repeatedly absorbed by more heavy nuclei

A) natural radioactivity commences
*B) a chain reaction commences
C) polymerization reactions occur
D) fusion reactions occur

13A.15.

29) Which of the following scientists is credited with the discovery of the formula for the equivalence of mass and energy?

A) Rutherford
B) Millikan
C) Maxwell
*D) Einstein

30) When an atom of $^{214}_{82}Pb$ emits a beta particle, the new element formed is

A) Tl
*B) Bi
C) Rn
D) Pt

31) The half life (the time required for half of a sample to decompose) of the isotope polonium-218 is 3 min. What fraction of the original sample of the polonium-218 remains after one hour?

*A) $\frac{1}{2^{20}}$

B) $\frac{1}{20^2}$

C) $\frac{1}{2 \times 20}$

D) $\frac{1}{20}$

13A.15.

32) When an atom of plutonium, Pu, decays, an alpha particle is produced along with a new atom of

A) Ra

*B) U

C) Th

D) Pa

33) Early in 1969, two isotopes of an artificial radioactive element number 104 were produced by bombarding a californium target with carbon nuclei in a heavy ion linear accelerator. If element 104 is given the symbol Ku, what is the missing species in the following equation?

$$^{249}_{98}Cf + ^{13}_{6}C \rightarrow ^{259}_{104}Ku + ----$$

A) $3\ ^{1}_{1}H$

B) $^{3}_{1}H$

*C) $3\ ^{1}_{0}n$

D) $^{4}_{2}He$

(c) S-17E Statistical Tables

ID #	GUIDELINE	DIFFICULTY	KEYWORDS
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13A.1.40	13A.1	**	1256
13A.1.41	13A.1		1256 750 979
13A.1.42	13A.1	**	701 1256 303
13A.1.43	13A.1	**	106 701 751 1256
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13A.1.91	13A.1	***	866 252 315
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13A.1.94	13A.1	***	1256 107

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13A.2.13	13A.2.1.e		411 311
13A.2.14	13A.2.3.b	***	1199 643
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13A.3.9	13A.3.2.d		214 438 971
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13A.3.11	13A.3.2.e		644 1138 1140 1155
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13A.3.15	13A.3.3.a	**	913 914 980
13A.3.16	13A.3.3.a		24 971 1140
13A.3.17	13A.3.3.a		24 434 531 1001
13A.3.18	13A.3.3.b	***	24 531 913 914
13A.3.19	13A.3.3.b	**	310 311 312 1080
13A.3.20	13A.3.1.	**	843 971 835
13A.3.21	13A.3.2.	*	544 971
13A.3.22	13A.3.2.a	*	971 979
13A.3.23	13A.3.2.a	**	971 1127
13A.3.24	13A.3.2.a	**	977 971 915
13A.3.25	13A.3.2.c		914 23
13A.3.26	13A.3.2.c	*	1259 971
13A.3.27	13A.3.3.b	**	311 914
13A.3.28	13A.3.3.b	*	914 434 1080
13A.3.29	13A.3.3.c	**	181 971 24
13A.3.30	13A.3.2.c	*	192
13A.3.34	13A.3.1.		971 972
13A.3.35	13A.3.2.e	***	1155 643
13A.3.36	13A.3.2.b	***	404 973
13A.3.37	13A.3.3.a	***	24 914
13A.3.38	13A.3.1.	***	1259 971
13A.3.39	13A.3.3.c	*	181 24 914
13A.3.40	13A.3.3.a	*	24 914
13A.3.41	13A.3.3.a	**	181 914 24
13A.3.42	13A.3.2.b	*	230
13A.3.43	13A.3.2.e	***	1155 643
13A.3.44	13A.3.3.a	**	914 23
13A.3.45	13A.3.3.b	***	914 311 531
13A.3.46	13A.3.3.b	*	405 412

ID #	GUIDELINE	DIFFICULTY	KEYWORDS
13A.4.1	13A.4.2.f	**	252 415 423 665 1140
13A.4.2	13A.4.2.f		134 665 861 1141
13A.4.3	13A.4.2.f	***	415 423 665 1139 1140
13A.4.4	13A.4.2.f	***	415 423 665
13A.4.5	13A.4.2.e		415 417 423 665
13A.4.6	13A.4.3.a	**	198 252 415 419 1140
13A.4.7	13A.4.3.a	*	210 419 421 662
13A.4.8	13A.4.2.d		415 423 665 918
13A.4.9	13A.4.2.d	**	198 415 423 665 1140
13A.4.10	13A.4.2.d	**	415 419 635 665 1140
13A.4.11	13A.4.3.d	**	307 412 927 962
13A.4.12	13A.4.3.e	*	415 419 420 635 637
13A.4.13	13A.4.3.e	**	252 415 419 635
13A.4.14	13A.4.3.e	***	252 415 417
13A.4.15	13A.4.3.e	***	415 416 417 420 635
13A.4.16	13A.4.2.d	*	1260 415
13A.4.17	13A.4.2.f	**	665
13A.4.18	13A.4.3.a	**	419 417
13A.4.19	13A.4.3.a	***	419 198
13A.4.20	13A.4.3.b	*	420 415
13A.4.21	13A.4.3.b	*	419 635
13A.4.22	13A.4.2.c	**	415 418
13A.4.23	13A.4.2.f	*	1001 415
13A.4.24	13A.4.3.e	**	415 417 1260
13A.4.25	13A.4.3.e	***	415 417 1260
13A.4.26	13A.4	**	695
13A.4.27	13A.4.2.b	***	351 415
13A.4.28	13A.4.2.d	**	1222 423 1260 15
13A.4.29	13A.4.2.f		415 665 198
13A.4.30	13A.4.2.f	**	665 415 1260
13A.4.31	13A.4.3.a	**	415 417 419 635
13A.4.32	13A.4.3.a	*	632 415 419
13A.4.33	13A.4.3.a	*	415 635 419 418
13A.4.34	13A.4.3.a	*	419 635 415 417
13A.4.35	13A.4.2.a	**	415 418 1095
13A.4.36	13A.4.2.a	***	11 971 983
13A.4.37	13A.4.3.e	*	419 415 417 635
13A.4.38	13A.4.3.e	***	415 417 591
13A.4.39	13A.4.3.e	**	635 415 417

ID #	GUIDELINE	DIFFICULTY	KEYWORDS
13A.5.1	13A.5.2.d	***	16 633 1104
13A.5.2	13A.5.1.d	***	1057
13A.5.3	13A.5.1.a		343 1054 1061
13A.5.4	13A.5.2.a	*	370 1107
13A.5.5	13A.5.2.b	**	76 370
13A.5.6	13A.5.2.b	*	363 370 1055
13A.5.7	13A.5.2.c	**	1054 1212
13A.5.8	13A.5.2.d	***	637 1025 1050
13A.5.9	13A.5.2.d	***	637 1055 1058
13A.5.10	13A.5.2.d	**	637 1020 1055
13A.5.11	13A.5.2.h	**	16 856
13A.5.12	13A.5.2.d	**	637 1054
13A.5.13	13A.5.2.e		637 1054 1055 1058
13A.5.14	13A.5.2.e	**	637 1054 1061
13A.5.15	13A.5.2.e	***	415 637 1054
13A.5.16	13A.5.2.e		417 637 1025 1054
13A.5.17	13A.5.2.f		76 915 1066
13A.5.18	13A.5.2.f	***	252 637 915 1054
13A.5.19	13A.5.2.h		180 415 637 1054 1055
13A.5.20	13A.5.2.h	***	637 1054 1058
13A.5.21	13A.5.2.b	***	1256 637 915 1054
13A.5.22	13A.5.2.d	*	637 1054 252 745
13A.5.23	13A.5.2.d	***	637 1025 1054 252
13A.5.24	13A.5.2.e		1054
13A.5.25	13A.5.2.g		1058 637 915
13A.5.26	13A.5.2.		341 106
13A.5.27	13A.5.2.		252 745
13A.5.28	13A.5.2.d		637 1054 637 1058
13A.5.29	13A.5.2.d		637 1058
13A.5.30	13A.5.2.d		637 1058 1054
13A.5.31	13A.5.2.e		1054 1025
13A.5.32	13A.5.2.d		1054 637 1058
13A.5.33	13A.5.2.		315 1059
13A.5.34	13A.5.2.		315 252 866
13A.5.35	13A.5.2.		252 866
13A.5.36	13A.5.2.a	**	1022 612
13A.5.37	13A.5.2.a	**	370 386
13A.5.38	13A.5.2.b	**	800
13A.5.39	13A.5.2.b	**	1022 613
13A.5.40	13A.5.2.d	**	1054 637 1058
13A.5.41	13A.5.2.g	***	915
13A.5.42	13A.5.2.h	**	1054 1025
13A.5.43	13A.5.2.h	*	866
13A.5.44	13A.5.2.h	**	252 747
13A.5.45	13A.5.2.h	***	637 1058 1054
13A.5.46	13A.5.2.h	***	252 747
13A.5.47	13A.5.2.f		915 1054
13A.5.48	13A.5.2.		1022 612
13A.5.49	13A.5.2.f		1054 1025
13A.5.50	13A.5.2.d		637 1058
13A.5.51	13A.5.2.		252 747
13A.5.52	13A.5.2.		370 609

ID #	GUIDELINE	DIFFICULTY	KEYWORDS
13A.5.53	13A.5.2.f		915
13A.5.54	13A.5.2.d		637 1058
13A.5.55	13A.5.2.		866 315 1059
13A.5.56	13A.5.2.e		1054
13A.5.57	13A.5.2.		866 1059
13A.5.58	13A.5.2.f		637 1058 915
13A.5.59	13A.5.3.a	***	222 11 1266 633
13A.5.60	13A.5.3.c	**	83 1262
13A.5.61	13A.5.3.a	**	341 11 1266

ID #	GUIDELINE	DIFFICULTY	KEYWORDS
13A.6.1	13A.6.1.	**	89 153 751
13A.6.2	13A.6.		16 111 751 1050
13A.6.3	13A.6.2.	**	16 75 76 1105 1106
13A.6.4	13A.6.3.a	**	16 252 1104 1222
13A.6.5	13A.6.1.b	**	16 415 610 638
13A.6.6	13A.6.3.c	**	524 531 748 880
13A.6.7	13A.6.1.c	***	16 252 568
13A.6.8	13A.6.2.d		16 562 566 787
13A.6.9	13A.6.2.d	**	16 787 1157
13A.6.10	13A.6.2.d	**	16 17 787 1157
13A.6.11	13A.6.2.d		16 787 1208
13A.6.12	13A.6.2.d	**	16 787 1208
13A.6.13	13A.6.2.d	**	16 585 881 1157
13A.6.14	13A.6.2.e	**	16 415 566 665 856
13A.6.15	13A.6.2.e	**	16 856 1266 1267
13A.6.16	13A.6.2.e	*	16 253 566 856
13A.6.17	13A.6.2.e	*	11 16 856 1061
13A.6.18	13A.6.2.e	**	16 856
13A.6.19	13A.6.3.a	**	11 16 415 419 633
13A.6.20	13A.6.3.a	**	16 633 1104
13A.6.21	13A.6.3.a	**	11 16 633 1105
13A.6.22	13A.6.3.a	***	16 462 633 856
13A.6.23	13A.6.3.a		11 16 252 562
13A.6.24	13A.6.3.a	***	16 251 341 633
13A.6.25	13A.6.3.a	**	16 234 419 562 585
13A.6.26	13A.6.3.b	***	16 158 159
13A.6.27	13A.6.3.b	**	16 157 159 933 1212
13A.6.28	13A.6.3.b	**	16 157 159 983 1061
13A.6.29	13A.6.3.b	**	16 11 156 158 603
13A.6.30	13A.6.2.e	*	11 856 252 562
13A.6.31	13A.6.2.e	**	11 856 252 562
13A.6.32	13A.6.2.d	**	787 252 562
13A.6.33	13A.6.2.b	**	83 111
13A.6.34	13A.6.1.b	**	638
13A.6.35	13A.6.1.b	***	620 330 11
13A.6.36	13A.6.3.b	*	111 16
13A.6.37	13A.6.2.e	**	856 252 562
13A.6.38	13A.6.3.a	***	16 585 17 787
13A.6.39	13A.6.3.a	**	585 620
13A.6.40	13A.6.3.a	**	633 419
13A.6.41	13A.6.3.b	**	17 585
13A.6.42	13A.6.3.c	***	159
13A.6.43	13A.6.3.a	***	856 252 633
13A.6.44	13A.6.2.e	*	856 252 562 567
13A.6.45	13A.6.3.b	**	15 620 562
13A.6.46	13A.6.3.c	**	111
13A.6.47	13A.6.2.d	***	17 787
13A.6.48	13A.6.3.b	***	856 562
13A.6.49	13A.6.3.a	*	62 633 341
13A.6.50	13A.6.3.c		11 83 159
13A.6.51	13A.6.3.a	**	633 111
13A.6.52	13A.6.2.d	*	17 787 252

ID #	GUIDELINE	DIFFICULTY	KEYWORDS
13A.6.53	13A.6.2.e	*	856 252 562
13A.6.54	13A.6.2.e	**	858 16
13A.6.55	13A.6.3.a	**	1222 16 342
13A.6.56	13A.6.3.c	**	16 159
13A.6.57	13A.6.3.a	**	633 856
13A.6.58	13A.6.2.b	**	16,1266,116
13A.6.59	13A.6.3.b	***	16 1222 1105
13A.6.60	13A.6.3.b	**	16 856
13A.6.61	13A.6.2.d	**	16 1105 1266
13A.6.62	13A.6.3.a	**	16 1266 1222
13A.6.63	13A.6.1.c	***	1266 638
13A.6.64	13A.6.3.a	***	1267 634 1223
13A.6.65	13A.6.2.e	*	16 856 1266
13A.6.66	13A.6.3.a	***	633 1266
13A.6.67	13A.6.1.a	*	370 670
13A.6.68	13A.6.3.b	***	370 670 1022
13A.6.69	13A.6.2.d	**	787 11 111
13A.6.70	13A.6.3.c	**	83 11
13A.6.71	13A.6.3.e	**	1106 11 261
13A.6.72	13A.6.2.d	**	1157 11 111 585
13A.6.73	13A.6.3.a	**	633 11
13A.6.74	13A.6.3.c	***	11 633 261
13A.6.75	13A.6.	***	796
13A.6.76	13A.6.2.e	**	1266 856
13A.6.77	13A.6.3.a	***	1267 633
13A.6.78	13A.6.2.	***	787
13A.6.79	13A.6.3.a	**	585 620 415 236
13A.6.80	13A.6.	*	1267 253
13A.6.81	13A.6.1.a	**	633 370 1266
13A.6.82	13A.6.1.a		1222 11 620
13A.6.83	13A.6.3.b	**	111
13A.6.84	13A.6.2.d	**	856 111 11 787
13A.6.85	13A.6.2.d	**	787 463
13A.6.86	13A.6.2.b	**	1267
13A.6.87	13A.6.2.e	***	856
13A.6.88	13A.6.2.b	*	1267 1212
13A.6.89	13A.6.3.c	***	83
13A.6.90	13A.6.1.b	*	1212 1224 787
13A.6.91	13A.6.1.	***	1267
13A.6.92	13A.6.2.	***	1266 1222
13A.6.93	13A.6.2.	***	252 11 315
13A.6.94	13A.6.3.a	***	633 11 1222
13A.6.95	13A.6.2.d	***	787 11 1105
13A.6.96	13A.6.2.d	**	787 11 1105 252
13A.6.97	13A.6.3.a	***	1266 633 11 1105
13A.6.98	13A.6.1.a	*	620 306
13A.6.99	13A.6.3.b	*	
13A.6.100	13A.6.2.	*	747 252
13A.6.101	13A.6.2.d	**	787
13A.6.102	13A.6.2.d	*	787 252 747
13A.6.103	13A.6.2.e	**	1266 856
13A.6.104	13A.6.3.a	**	633 856

ID #	GUIDELINE	DIFFICULTY	KEYWORDS
13A.6.105	13A.6.3.b	***	856 11 1222
13A.6.106	13A.6.3.b	***	11 633 1105 1222
13A.6.107	13A.6.3.c	**	159 16 261
13A.6.108	13A.6.2.b	***	16 111 585
13A.6.109	13A.6.3.a	***	16 633 1222 1266
13A.6.110	13A.6.	***	16 253 1105
13A.6.111	13A.6.1.c	**	16 787
13A.6.112	13A.6.3.b	*	16 585
13A.6.113	13A.6.2.d	***	16 1106 1266 1267
13A.6.114	13A.6.3.b	**	16 585
13A.6.115	13A.6.2.e	**	16 856 1266
13A.6.116	13A.6.1.c	**	16 1266 1106
13A.6.117	13A.6.1.c	***	16 1266 1106
13A.6.118	13A.6.3.b	***	16 856 1106 1222
13A.6.119	13A.6.	**	252 16
13A.6.120	13A.6.2.e		856 16 1266
13A.6.121	13A.6.	***	16 1256
13A.6.122	13A.6.3.a		16 633 1222 1266
13A.6.123	13A.6.3.a	***	16 17 1222 1106
13A.6.124	13A.6.2.e	**	16 856 1266
13A.6.125	13A.6.3.a	***	16 633 1222
13A.6.126	13A.6.2.d	***	16 17 252 1105 1106
13A.6.127	13A.6.3.a	***	16 633 856 1222
13A.6.128	13A.6.1.c	*	16 787
13A.6.129	13A.6.1.c	***	16 1266 1106
13A.6.130	13A.6.2.d	***	746 568 252 858
13A.6.131	13A.6.1.a	**	341 780 68
13A.6.132	13A.6.2.a	***	16 20 1105
13A.6.133	13A.6.3.a	***	16 33 856
13A.6.134	13A.6.2.d	***	16 787
13A.6.135	13A.6.3.b	***	16 341
13A.6.136	13A.6.2.e	**	16 856 1267
13A.6.137	13A.6.3.	***	16 1105 1106 341
13A.6.138	13A.6.2.d	***	16 20 17
13A.6.139	13A.6.2.e	**	16 856 785
13A.6.140	13A.6.3.c	***	16 83
13A.6.141	13A.6.3.b	***	16 158 159
13A.6.142	13A.6.2.d	***	16 17 1105 1106
13A.6.143	13A.6.3.a		10 11 16 252 633
13A.6.144	13A.6.3.b		16 159 620

ID #	GUIDELINE	DIFFICULTY	KEYWORDS
13A.7.1	13A.7.1.b	***	353 509 988
13A.7.2	13A.7.2.b	**	353 371 911 988
13A.7.3	13A.7.2.b	***	353 512 988
13A.7.4	13A.7.2.b	**	353 509 988
13A.7.5	13A.7.2.b		353 510 988 1088
13A.7.6	13A.7.2.b		242 353 367
13A.7.7	13A.7.2.b		353 988 1206 1207
13A.7.8	13A.7.2.b	***	353 843 849 988 993
13A.7.9	13A.7.2.c	*	850 980 988
13A.7.10	13A.7.2.c		843 988 992
13A.7.11	13A.7.2.c	*	843 847 988 990
13A.7.12	13A.7.2.c	**	699 844 988
13A.7.13	13A.7.2.c	**	844 988
13A.7.14	13A.7.2.c	**	844 988
13A.7.15	13A.7.2.c	***	844 988
13A.7.16	13A.7.2.c		843 844 988 992
13A.7.17	13A.7.2.c	**	844 988
13A.7.18	13A.7.2.c	*	844 988
13A.7.19	13A.7.2.c		844 988
13A.7.20	13A.7.2.c	***	843 849 988 990
13A.7.21	13A.7.2.c	**	840 850 988 992
13A.7.22	13A.7.2.c		844 850 988 992
13A.7.23	13A.7.3.a	**	509 511 988
13A.7.24	13A.7.3.a	**	509 511 988
13A.7.25	13A.7.3.a	**	511 845 988
13A.7.26	13A.7.3.a	***	511 845 988
13A.7.27	13A.7.3.a		107 511 845 989
13A.7.28	13A.7.3.b	***	107 751 845 988 992
13A.7.29	13A.7.4.a		369 751 988 992
13A.7.30	13A.7.4.a	*	70 184 368 371
13A.7.31	13A.7.2.a	**	512 1206 1207 367
13A.7.32	13A.7.2.b	***	512 367 1206
13A.7.33	13A.7.3.a	**	989 386 367
13A.7.34	13A.7.4.b	**	367 369 276
13A.7.35	13A.7.1.b	**	989 386
13A.7.36	13A.7.2.c	***	844
13A.7.37	13A.7.1.b	***	850 989
13A.7.38	13A.7.3.a	*	107 511 367
13A.7.39	13A.7.2.c	***	567 846
13A.7.40	13A.7.1.b	***	850 993 988
13A.7.41	13A.7.2.b	**	1207 843 367 368
13A.7.42	13A.7.4.b	**	1207 843 367
13A.7.43	13A.7.2.c	*	1256 992 843 988
13A.7.44	13A.7.2.c	**	844 988
13A.7.45	13A.7.2.a	**	843 512 1088 988
13A.7.46	13A.7.2.b	*	353 988
13A.7.47	13A.7.3.a	***	511 988
13A.7.48	13A.7.	*	988 1086
13A.7.49	13A.7.3.b	**	845 511 988
13A.7.50	13A.7.1.b	***	988 990
13A.7.51	13A.7.3.b	**	988 845 511
13A.7.52	13A.7.1.b		988 509

ID #	GUIDELINE	DIFFICULTY	KEYWORDS
13A.7.53	13A.7.1.b	**	988 992 843
13A.7.54	13A.7.4.b		369 751 376
13A.7.55	13A.7.2.c		249 844 394
13A.7.56	13A.7.2.a	***	74 605 381
13A.7.57	13A.7.2.c	**	846
13A.7.58	13A.7.2.c	*	844
13A.7.59	13A.7.2.c	**	846
13A.7.60	13A.7.2.a	**	369 367 509 1206
13A.7.61	13A.7.2.a	***	990
13A.7.62	13A.7.2.a	**	843 509
13A.7.63	13A.7.2.b	**	509 1082
13A.7.64	13A.7.2.a	**	993 386 988
13A.7.65	13A.7.3.		1256 107
13A.7.66	13A.7.2.b	***	371 367 368 609
13A.7.67	13A.7.3.	**	1256 107
13A.7.68	13A.7.2.c	***	850
13A.7.69	13A.7.2.a	***	988 990
13A.7.70	13A.7.2.c	*	844
13A.7.71	13A.7.3.	**	1256 107
13A.7.72	13A.7.3.	**	1256 107
13A.7.73	13A.7.3.	***	989 1256 107
13A.7.74	13A.7.3.	*	1256 107
13A.7.75	13A.7.2.c	*	796
13A.7.76	13A.7.2.a	**	991 990
13A.7.77	13A.7.2.b	**	510 391 367
13A.7.78	13A.7.3.a	*	107 1256 509
13A.7.79	13A.7.2.c	*	844
13A.7.80	13A.7.2.a	**	26 982 99

ID #	GUIDELINE	DIFFICULTY	KEYWORDS
13A.8.1	12A.8.2.c	*	252 747
13A.8.2	12A.8.2.c	***	343 866 1059 1061
13A.8.3	13A.8.1.a	***	401 538 1041
13A.8.4	13A.8.1.a		91 195 401 438 463
13A.8.5	13A.8.1.b		303 414 484 751 1208
13A.8.6	13A.8.1.c		59 97 276 447 988
13A.8.7	13A.8.2.a		185 362 697
13A.8.8	13A.8.2.b	**	678 704 706 995 1075
13A.8.9	13A.8.2.e	*	98 803 1009 1010
13A.9.1	13A.9.1.a		374 470 673 947
13A.9.2	13A.9.1.a	**	388 470 1076 1221
13A.9.3	13A.9.1.a	**	374 574 955 1180 1228
13A.9.4	13A.9.1.d	**	130 131 948
13A.9.5	13A.9.1.e		388 408 647 949
13A.9.6	13A.9.1.f	**	376 829 947
13A.9.7	13A.9.1.h		380 949
13A.9.8	13A.9.2.b		388 1193
13A.9.9	13A.9.2.b	***	101 388 832 949 1114
13A.9.10	13A.9.2.b		101 388 832
13A.9.11	13A.9.2.b	**	388 395 501 874 1193
13A.9.12	13A.9.2.h	**	98 388 775 862 921
13A.9.13	13A.9.3.b		388 501 621
13A.9.14	13A.9.3.a	*	388 621
13A.9.15	13A.9.3.b		387 621 874 986
13A.9.16	13A.9.3.b		621 874
13A.9.17	13A.9.3.b	***	388 605
13A.9.18	13A.9.3.b	**	388 605
13A.9.19	13A.9.1.e	***	698 829
13A.9.20	13A.9.2.b	**	382 94
13A.9.21	13A.9.2.b	***	829 1195 501
13A.9.22	13A.9.2.b	*	39 445 853
13A.9.23	13A.9.3.b	**	620 794
13A.9.24	13A.9.3.b	**	38 620
13A.9.25	13A.9.	**	605 703
13A.9.26	13A.9.2.b		1250
13A.9.27	13A.9.2.b	**	703 1250
13A.9.28	13A.9.2.b	*	722 605 910
13A.9.29	13A.9.2.b	***	1195 501
13A.9.30	13A.9.2.a	***	89 501 949 376
13A.9.31	13A.9.2.b	***	829 376
13A.9.32	13A.9.3.c	**	847 445 844
13A.9.33	13A.9.2.b	*	388 501 829
13A.9.34	13A.9.1.a	**	374 893 405 1221
13A.9.35	13A.9.3.c	***	501 619
13A.9.36	13A.9.3.c	***	621 405 408
13A.9.37	13A.9.2.b	*	1195 388
13A.9.38	13A.9.1.e	***	949
13A.9.39	13A.9.1.f	***	949 829
13A.9.40	13A.9.	***	
13A.9.41	13A.9.	***	
13A.9.42	13A.9.1.g	**	670 388 1233 382

ID #	GUIDELINE	DIFFICULTY	KEYWORDS
13A.9.43	13A.9.1.g	*	670 382 388 1149
13A.9.44	13A.9.1.g	**	670 382 388 1149
13A.9.45	13A.9.1.d	***	408 558 673 470
13A.9.46	13A.9.1.g	*	670 388 382
13A.9.47	13A.9.2.b	**	445 872 874
13A.9.48	13A.9.3.a	***	378
13A.9.49	13A.9.3.c	**	621 806
13A.9.50	13A.9.3.c	**	1250 501 38
13A.9.51	13A.9.1.h	***	949 647
13A.9.52	13A.9.1.g	*	670 382 853 358
13A.9.53	13A.9.3.c	**	621 874
13A.9.54	13A.9.1.e	**	949 775 558
13A.9.55	13A.9.1.f	*	297
13A.9.56	13A.9.1.h		949 775 921
13A.9.57	13A.9.3.b	**	200 874
13A.9.58	13A.9.3.a	***	151
13A.9.59	13A.9.2.b	**	1250 801
13A.9.60	13A.9.2.b	*	1250 94
13A.9.61	13A.9.3.b	***	621 874
13A.9.62	13A.9.1.e	**	949 647 693
13A.9.63	13A.9.2.b		794 388
13A.9.64	13A.9.2.b		1250 949
13A.9.65	13A.9.1.h		408 949 829
13A.9.66	13A.9.1.h		1250 727 949 1083 605
13A.9.67	13A.9.1.f		1195 855
13A.9.68	13A.9.1.f		829 922
13A.9.69	13A.9.1.h		829 949 408
13A.9.70	13A.9.2.a	*	829 405
13A.9.71	13A.9.1.f		855 379 829
13A.9.72	13A.9.2.b		388 433 1250
13A.9.73	13A.9.1.a	**	1220 673 888
13A.9.74	13A.9.3.b	**	384 621 874
13A.9.75	13A.9.1.h	*	829 408
13A.9.76	13A.9.2.b	***	234
13A.9.77	13A.9.1.f	*	1250 94 380
13A.9.78	13A.9.3.c	***	621
13A.9.79	13A.9.1.d	***	408 433 1164
13A.9.80	13A.9.3.b	**	621 874 1167
13A.9.81	13A.9.1.d	**	397 894 470 888 433
13A.9.82	13A.9.1.c	**	829 501 946 949
13A.9.83	13A.9.3.c	***	621 1250
13A.9.84	13A.9.2.a	*	408 501 1250
13A.9.85	13A.9.3.c	***	621 1250
13A.9.86	13A.9.1.	***	1029 1220
13A.9.87	13A.9.2.b	*	445 1250 949 380 388
13A.9.88	13A.9.2.b	*	1250 380 388 395 148
13A.9.89	13A.9.3.b	*	873 874 722 621
13A.9.90	13A.9.2.b	*	1250 647 775 949 605
13A.9.91	13A.9.1.f	*	829 831 832 922
13A.9.92	13A.9.2.b	*	433 1250 832
13A.9.93	13A.9.2.b	*	1250 775 949 647
13A.9.94	13A.9.3.b	***	621 405 874 806
13A.9.95	13A.9.3.b	*	384 621 415 874

ID #	GUIDELINE	DIFFICULTY	KEYWORDS
13A.10.1	13A.10.4.a	**	753 1069 1144 1145
13A.10.2	13A.10.2.	**	611 147 388
13A.10.3	13A.10.2.		621 617
13A.10.4	13A.10.2.b	**	784
13A.10.5	13A.10.2.b		341 754
13A.10.6	13A.10.1.c		1193 38
13A.10.7	13A.10.3.	**	147 278 611 836 1197
13A.10.8	13A.10.1.a	***	147 206
13A.10.9	13A.10.2.	**	382 670
13A.10.10	13A.10.2.	*	382 670
13A.10.11	13A.10.2.a	*	149 463 1193
13A.10.12	13A.10.2.a	*	149 1192
13A.10.13	13A.10.2.a	**	3
13A.10.14	13A.10.3.	**	147 611 621
13A.10.15	13A.10.3.		147 278 1174 1197
13A.10.16	13A.10.4.	***	802 900
13A.10.17	13A.10.4.	**	147 753 802 898 900
13A.10.18	13A.10.4.	***	753 1037
13A.10.19	13A.10.3.b	*	143 612
13A.10.20	13A.10.1.a	**	143 408
13A.10.21	13A.10.3.b	**	898
13A.10.22	13A.10.3.b	**	280 249
13A.10.23	13A.10.2.	*	38 829
13A.10.24	13A.10.5.	*	279
13A.10.25	13A.10.4.a	**	753
13A.10.26	13A.10.3.b	**	370
13A.10.27	13A.10.4.b	**	333
13A.10.28	13A.10.4.a	*	753
13A.10.29	13A.10.4.a	**	753
13A.10.30	13A.10.4.a	*	753
13A.10.31	13A.10.4.b	***	333 898
13A.10.32	13A.10.3.b	**	279
13A.10.33	13A.10.3.b	**	611
13A.10.34	13A.10.5.	*	346
13A.10.35	13A.10.4.a	*	753
13A.10.36	13A.10.2.	**	551 147
13A.10.37	13A.10.1.a	***	599 460
13A.10.38	13A.10.	*	670 382
13A.10.39	13A.10.3.b	***	279 280
13A.10.40	13A.10.2.	***	279
13A.10.41	13A.10.1.b	***	147 1197 680
13A.10.42	12A.10.4.b	**	626 1109 221
13A.10.43	13A.10.1.c	*	900 463
13A.10.44	13A.10.4.a	***	1037 753
13A.10.45	13A.10.4.a	*	1037 753
13A.10.46	13A.10.4.b	**	1037 753 56
13A.10.47	13A.10.4.a	*	1037 753 172
13A.10.48	13A.10.4.a	**	753 1144
13A.10.49	13A.10.4.a	*	753 1144
13A.10.50	13A.10.4.b	**	559 560 1212
13A.10.51	13A.10.4.b	**	900 278 559 147
13A.10.52	13A.10.3.	**	899 279

ID #	GUIDELINE	DIFFICULTY	KEYWORDS
13A.10.53	13A.10.3.c	**	899 279
13A.10.54	13A.10.3.c	**	900 901
13A.10.55	13A.10.4.a	**	137 686 999
13A.10.56	13A.10.4.a	***	147 551 759
13A.10.57	13A.10.4.a	**	1037 1112 1144 1212
13A.10.58	13A.10.2.	**	149 853 1193
13A.10.59	13A.10.2.	**	280 147 249 1268
13A.10.60	13A.10.4.a	***	149 753 332 249
13A.10.61	13A.10.1.c	***	279 147
13A.10.62	13A.10.2.	**	832 873 611 41
13A.10.63	13A.10.4.a	***	332 900
13A.10.64	13A.10.1.c	**	147 149 832
13A.10.65	13A.10.3.	***	898 621
13A.10.66	13A.10.4.b	**	898 900 332
13A.10.67	13A.10.3.b	***	612 621
13A.10.68	13A.10.2.	*	149 463 1110
13A.10.69	13A.10.4.a	***	753 820
13A.10.70	13A.10.4.b	***	900
13A.10.71	13A.10.4.a	***	829 1069 1068 753
13A.10.72	13A.10.4.a	**	753 1109 1186
13A.10.73	13A.10.2.	***	1193 1250 1068 401
13A.10.74	13A.10.4.	***	753 621 900
13A.10.75	13A.10.2.	***	463 1041 1068

ID #	GUIDELINE	DIFFICULTY	KEYWORDS
13A.11.1	13A.11.1.a	***	147 279 559 611 1197
13A.11.2	13A.11.2.d	*	617
13A.11.3	13A.11.2.b		617 761 801 1152
13A.11.4	13A.11.1.a	**	1197
13A.11.5	13A.11.2.f	*	597 135
13A.11.6	13A.11.2.a	**	1197
13A.11.7	13A.11.2.a	**	1197
13A.11.8	13A.11.2.a	*	1197
13A.11.9	13A.11.1.b	**	614
13A.11.10	13A.11.2.d	**	1194 1115
13A.11.11	13A.11.1.b	**	1194
13A.11.12	13A.11.2.e	**	341 900
13A.11.13	13A.11.2.d	**	613 900
13A.11.14	13A.11.2.e	**	900
13A.11.15	13A.11.1.c	***	722 380
13A.11.16	13A.11.1.a	**	1197 761
13A.11.17	13A.11.1.a	***	1197 597 445
13A.11.18	13A.11.2.f	***	560
13A.11.19	13A.11.2.d	***	652 404
13A.11.20	13A.11.2.b	**	279 258
13A.11.21	13A.11.1.b	**	279
13A.11.22	13A.11.2.d	**	387 611
13A.11.23	13A.11.2.b	**	784 518
13A.11.24	13A.11.2.d	**	611
13A.11.25	13A.11.2.b	**	784
13A.11.26	13A.11.2.d	***	611
13A.11.27	13A.11.2.f	***	559
13A.11.28	13A.11.2.a	***	761 784
13A.11.29	13A.11.1.a		1197
13A.11.30	13A.11.2.a	**	559
13A.11.31	13A.11.2.d	**	614
13A.11.32	13A.11.2.a		754
13A.11.33	13A.11.2.f	***	559
13A.11.34	13A.11.2.f	**	1212 559 135 716
13A.11.35	13A.11.2.e	*	279 147 280 801 387
13A.11.36	13A.11.2.f	*	559 147
13A.11.37	13A.11.2.d	*	802 784 258
13A.11.38	13A.11.2.a	***	802 278 1197 147
13A.11.39	13A.11.2.d	**	147 802 1197
13A.11.40	13A.11.1.c	**	725 147
13A.11.41	13A.11.2.e	**	1054 1060
13A.11.42	13A.11.2.	**	333 900
13A.11.43	13A.11.2.e	**	898 900
13A.11.44	13A.11.2.	*	1197 147
13A.11.45	13A.11.2.e	**	900 1069 1060
13A.11.46	13A.11.2.b	***	754
13A.11.47	13A.11.1.a	**	1197 147
13A.11.48	13A.11.2.a	***	802 147
13A.11.49	13A.11.2.d		611 617 147
13A.11.50	13A.11.2.b	***	784
13A.11.51	13A.11.2.	*	899
13A.11.52	13A.11.1.c		722 147

ID #	GUIDELINE	DIFFICULTY	KEYWORDS
13A.11.53	13A.11.2.e	**	900 1109
13A.11.54	13A.11.1.c	***	611 614 725 724 285
13A.11.55	13A.11.2.e	**	1212 823 147
13A.11.56	13A.11.2.a	*	1237 597
13A.11.57	13A.11.2.a	***	1196
13A.11.58	13A.11.1.b	**	832 873 41 784
13A.11.59	13A.11.2.a	**	147 484 1197 1053
13A.11.60	13A.11.2.	**	484 147 1096 1199 1237
13A.11.61	13A.11.	**	597 835 901 1054 1064
13A.11.62	13A.11.1.c	***	272 285 288 778
13A.11.63	13A.11.2.b	*	288 41 282
13A.11.64	13A.11.2.f	*	41 559
13A.11.65	13A.11.2.b	**	784 1053
13A.11.66	13A.11.2.f	**	559
13A.11.67	13A.11.1.c	***	832 873 725 41
13A.11.68	13A.11.1.c		832 873 611 41
13A.11.69	13A.11.2.	*	518 147 279
13A.11.70	13A.11.2.f	**	41 135 559
13A.11.71	13A.11.2.b	**	784 147 753 1112
13A.11.72	13A.11.2.a	***	258 285 754
13A.11.73	13A.11.1.c	**	716 147 279
13A.11.74	13A.11.	**	41 206 377 378
13A.11.75	13A.11.2.e	**	900 761 1053
13A.11.76	13A.11.2.b	**	1053 41 725
13A.11.77	13A.11.2.e	**	900 332 333 901 898
13A.11.78	13A.11.2.a	***	1197 597 147
13A.11.79	13A.11.2.d	**	285 553 518
13A.11.80	13A.11.1.c	**	859 287 288
13A.11.81	13A.11.1.	*	206 100 143 147
13A.11.82	13A.11.2.	**	611 612 219
13A.11.83	13A.11.1.b	**	143 406 147 150
13A.11.84	13A.11.1.c	**	722 928 801 1193 874
13A.11.85	13A.11.2.b	**	784 30 281 282
13A.11.86	13A.11.1.a	***	874 1197 515 147 597

ID #	GUIDELINE	DIFFICULTY	KEYWORDS
13A.12.1	13A.12.	**	248 401 835 1041
13A.12.2	13A.12.2.a	**	556 626 627 825 1099
13A.12.3	13A.12.3.b	**	240 414 833 835
13A.12.4	13A.12.3.c	*	33 476 835 836
13A.12.5	13A.12.3.c	*	476 536 796 835 952
13A.12.6	13A.12.5.b		33 564
13A.12.7	13A.12.2.a	**	1041
13A.12.8	13A.12.3.e	***	426 871
13A.12.9	13A.12.2.a	***	626
13A.12.10	13A.12.2.a	*	1041
13A.12.11	13A.12.3.c	***	34 476
13A.12.12	13A.12.4.	**	796
13A.12.13	13A.12.4.	**	626 796
13A.12.14	13A.12.6.a	***	906
13A.12.15	13A.12.4.	**	796
13A.12.16	13A.12.3.e	***	426 16
13A.12.17	13A.12.5.b	***	796 1188
13A.12.18	13A.12.3.b	*	107 1256
13A.12.19	13A.12.3.e	*	426 1133
13A.12.20	13A.12.4.	**	954 796
13A.12.21	13A.12.3.e	***	51 796 476
13A.12.22	13A.12.3.b	***	796 843 640
13A.12.23	13A.12.4.	*	626
13A.12.24	13A.12.4.	**	796 33
13A.12.25	13A.12.4.	**	427 796
13A.12.26	13A.12.3.b	***	843 1188
13A.12.27	13A.12.3.c	**	11 1222
13A.12.28	13A.12.3.	***	834 835 928
13A.12.29	13A.12.3.c	***	1109 463 640 835
13A.12.30	13A.12.6.a		906 835
13A.12.31	13A.12.2.b	***	1143 979 990 835
13A.12.32	13A.12.4.	***	796 426 835
13A.12.33	13A.12.2.a	**	551 1069 169 147 928
13A.12.34	13A.12.5.b	***	556 1188 796 835
13A.12.35	13A.12.4.	***	834 835 796
13A.12.36	13A.12.5.b	**	835 836 627 221
13A.12.37	13A.12.3.e	***	835 426
13A.12.38	13A.12.3.c	**	835 640
13A.12.39	13A.12.3.	*	834 835 928
13A.12.40	13A.12.3.e	**	426 33 796 835
13A.12.41	13A.12.3.	**	476 51 796
13A.12.42	13A.12.4.	**	296 835 494
13A.12.43	13A.12.5.	***	402 1047 739
13A.12.44	13A.12.3.e	*	564 1212 426 835
13A.12.45	13A.12.6.c	*	865 53 930 143
13A.12.46	13A.12.4.	**	756 796 835 626 576
13A.12.47	13A.12.4.	*	626 221 835
13A.12.48	13A.12.3.c	***	796 835 34
13A.12.49	13A.12.2.a	*	866 757
13A.12.50	13A.12.4.	*	796 835
13A.12.51	13A.12.2.a	***	796 835 33
13A.12.52	13A.12.2.a		626 756 835

ID #	GUIDELINE	DIFFICULTY	KEYWORDS
13A.12.53	13A.12.3.e	**	835 181 564
13A.12.54	13A.12.3.	**	834 835 928
13A.12.55	13A.12.2.a	***	756 835 463 796
13A.12.56	13A.12.5.a	**	1236 835 248
13A.12.57	13A.12.5.b	**	137 835 1174
13A.12.58	13A.12.5.a	***	626 835 796
13A.12.59	13A.12.4.	*	796 835 576
13A.12.60	13A.12.4.	**	796 835 640

ID #	GUIDELINE	DIFFICULTY	KEYWORDS
13A.14.1	13A.14.2.a	***	236 34
13A.15.1	13A.15.2.c	***	535 320
13A.15.2	13A.15.2.d	**	49 122 957
13A.15.3	13A.15.2.d	**	957 122
13A.15.4	13A.15.2.d	*	811 49
13A.15.5	13A.15.2.d	*	453 811
13A.15.6	13A.15.2.c	*	49 811
13A.15.7	13A.15.2.c	***	122 1083 513
13A.15.8	13A.15.2.d	***	959 811
13A.15.9	13A.15.2.d	*	956
13A.15.10	13A.15.2.d		959
13A.15.11	13A.15.2.d	**	49
13A.15.12	13A.15.2.	*	513 956
13A.15.13	13A.15.2.c	**	813 49
13A.15.14	13A.15.2.c	***	813 122
13A.15.15	13A.15.2.c	*	956
13A.15.16	13A.15.2.d	**	956 513
13A.15.17	13A.15.2.c	**	808
13A.15.18	13A.15.2.c	**	122 49 956
13A.15.19	13A.15.2.d	***	1083 122 513
13A.15.20	13A.15.2.d	**	122 957
13A.15.21	13A.15.2.a	***	956 873
13A.15.22	13A.15.2.e	**	354 811
13A.15.23	13A.15.2.c	*	122 620 811 813 1255
13A.15.24	13A.15.2.c	**	513 628 956
13A.15.25	13A.15.2.c	*	628 788 808
13A.15.26	13A.15.2.e	**	812 953 809
13A.15.27	13A.15.2.d	*	49 957
13A.15.28	13A.15.2.e	**	453 192
13A.15.29	13A.15.2.e	*	542 405 701
13A.15.30	13A.15.2.c	*	122 809 811 957
13A.15.31	13A.15.2.c	**	513 302
13A.15.32	13A.15.2.c	*	957 49
13A.15.33	13A.15.2.	*	956 959 480

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